

Set 1: Significant figures and unit conversions

Exponential notation

This notation, also known as scientific notation, is used for convenience when writing very large or very small numbers.

To express a number using experimental notation, it is written as a number between 1 and 10 multiplied by the appropriate power of 10.

For example 0.056 expressed in exponential notation is written as 5.6×10^{-2} while 167,000 would be written as 1.67×10^5 .

Arithmetic using exponential notation is governed by the following rules:

Addition and subtraction	Multiplication	Division
Before adding or subtracting, numbers must be expressed to the same powers of 10.	When multiplying powers of 10 add their indices algebraically.	When dividing powers of 10 by each other, subtract the index of the denominator from that of the numerator.
Example: $(2.04 \times 10^5) + (4.7 \times 10^4)$ $= (2.04 \times 10^5) + (0.47 \times 10^5)$ $= (2.04 + 0.47) \times 10^5$ $= 2.51 \times 10^5$	Example: $(5 \times 10^5) \times (4 \times 10^2)$ $= (5 \times 4) \times 10^{5+2}$ $= 20 \times 10^7$ $= 2 \times 10^8$	Examples: $\frac{1.8 \times 10^8}{6 \times 10^5} = \frac{1.8}{6} \times 10^{8-5}$ $= 0.3 \times 10^3$ $= 3 \times 10^2$

Notes

Set 1: Exercises

- How many significant figures are there in the following?
 - 123
 - 1.23
 - 1.23×10^5
 - 1.23×10^{-59}
 - 123 000
 - 120 300
 - 102 003
 - 0.00000123
 - Express each of the following in scientific notation.
 - 6 409
 - 0.032
 - 891 000
 - 53.8
 - 0.0000061
 - Complete the following:
 - $8 \times 10^{-4} \text{ m} = \text{ mm}$
 - $4.5 \times 10^3 \text{ g} = \text{ kg}$
 - $9.0 \times 10^{-2} \text{ kg} = \text{ g}$
 - $7.03 \times 10^5 \text{ mL} = \text{ L}$
 - $0.05 \times 10^4 \text{ L} = \text{ mL}$
 - $2.59 \text{ nm} = \text{ m}$
 - A mixture is prepared using 3.104 g of substance A, 0.72 g of B, 16.2 g of D, and 0.002 g of E. What is the total mass of the mixture, to the correct number of significant figures, assuming no losses occur?

Notes

5. An atom of sodium weighs 3.819×10^{-23} g. How many sodium atoms are there in 20 kg of sodium?
6. Complete the following conversion table for pressure using scientific notation:

	pressure/mmHg	pressure/atm	pressure/Pa
a	760	1.00	1.01×10^5
b	750		
c		2.05	
d	100		
e			7.31×10^3

7. A synthetic mixture was prepared from compounds abundant in nature. The constituents were 0.103 g of CaCO₃, 11.45 g of Fe₂O₃, 0.01 g of NaCl, 0.001 g of KCl, 68.53 g of SiO₂. What would be the total mass, to the correct number of significant figures, of the mixture?

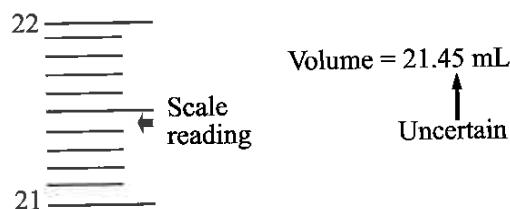
Set 2: Errors

Uncertainty and significant figures in measurement

When handling experimental data it is important to consider the appropriate number of significant figures to use and not be tempted to record values with large numbers of significant figures that may be produced by your calculator when processing data. Significant figures in a number are those digits that are known with certainty plus the first digit that is uncertain. All measurements have some level of uncertainty. These arise from the quality of the equipment used, the type of scale and the skill of the experimenter. Some apparatus will have a manufacturers' uncertainty assigned; however, there are cases where you will make a judgement.

Burettes are used in titration experiments to accurately measure the volume of a reactant added to another. A typical 50 mL burette has an analogue scale with a smallest scale division of 0.1 mL. A manufacturer's uncertainty of ± 0.05 mL (representing half the smallest scale division) is often assigned. Depending on your skill level you may judge that you can read the scale on the burette to the nearest 0.02 mL and so reduce the range of uncertainty for each measurement to ± 0.02 mL.

The burette reading in the diagram of 21.45 mL is estimated to the nearest 0.05 mL. The true value lies in a range from 21.40–21.50 mL.



Significant figures in a number are those digits that are known with certainty plus the first digit that is uncertain. In the reading of 21.45 mL the 2, 1 and 4 are known with certainty and the 5 is uncertain. This value has 4 significant figures.

	Number	Number of significant figures
1.	713	3
2.	7.03×10^3	3
3.	11.05	4
4.	0.027	2
5.	9.9643×10^{-7}	5

Whole numbers ending with one or more zeros such as 430 and 500 are ambiguous with respect to the number of significant figures because it is unclear whether the terminating zeros are significant or merely serve to locate the decimal point.

The use of scientific notation avoids ambiguity as shown by the following examples:

1. If 430 has 3 significant figures it is written as 4.30×10^2
2. If 430 has 2 significant figures it is written as 4.3×10^2
3. If 500 has 1 significant figure it is written as 5×10^2

Rounding off

Rounding is used to reduce the complexity of a number when it is written with more digits than are wanted or justified. The last digit written should give the best approximation of the number as it was before rounding. If the number before rounding is as close to one number as another, the one ending with an even digit is chosen, zero being regarded as even (example 6).

The following examples illustrate rounding to **three significant figures**.

1.	1.294	rounds to	1.29
2.	8.12349	rounds to	8.12
3.	0.01249	rounds to	0.0125
4.	18.951	rounds to	19.0
5.	7.1451	rounds to	7.15
6.	7.145	rounds to	7.14

Rounding in multi-step calculations

When performing calculations requiring several steps, only round to the appropriate number of significant figures after the final step. This avoids possible errors that can accumulate during a calculation if rounding to the strict number of significant figures is carried out at each step.

Set 2: Exercises

- An analyst is asked to find the iron content in a waste water sample. She chooses to use a four step method.
 - weigh a sample of waste water
 - precipitate the soluble iron by using hydroxide ions,
 - filter the sample and then heat the filtrate until all the iron hydroxide is converted to iron (III) oxide and the ash-less filter paper burns away.
 - weigh the iron (III) oxide sample

Steps 1 and 4 involve two uses of a balance - finding before and after weights. Her balance is capable of an accuracy of ± 1 mg. The mass of her original sample was 12.363 g and that of her iron (III) oxide residue was 0.834 g.

- What was the percentage uncertainty of each of the two masses recorded?
- What is the percentage uncertainty of the final result?

- An investigator reported the volume of liquid escaping from a factory as $36\ 671 \pm 153$ L. What is the percentage uncertainty of this result?
- A balance indicates masses correct to ± 0.2 mg. What is the minimum mass of a sample which must be taken if the weighing error is not to exceed
 - 1 part in 3000?
 - 0.01%?
- If a piece of plastic has a mass which may be as little as 80.1 g or as much as 80.5 g. If it is cut into two pieces, one of which weighs 40.0 ± 0.2 g.
 - What is the smallest mass that the other piece could have?
 - What is the mass of this piece, expressed in a plus or minus notation?

Set 2: Errors

Notes

5. Divide 6.8245 by 1.13, expressing your answer to the correct number of significant figures.
6. In processing the results of an experiment a student's final calculation is this

$$\frac{0.574862 \times 100}{156.0 \times 16.1}$$

What answer should he write down?

7. Convert the following Fahrenheit temperatures to Celsius readings:(To convert Fahrenheit to Celsius use the formula ${}^{\circ}\text{C} = ({}^{\circ}\text{F} - 32) \times 5/9$)
 - (a) 332°
 - (b) 1°
8. These data show the first ionisation energies of various atoms measured in the unit electronvolt. Convert these values to kJ mol^{-1} ($1 \text{ eV} = 96.485 \text{ kJ mol}^{-1}$).
 - (a) H:13.6
 - (b) He:24.6
 - (c) Be:9.32
9. Using the radioactive dating method, the linen wrapping of one of the Dead Sea Scrolls was found to be 1920 ± 350 years old. What is the percentage uncertainty of this result?
10. In solving a chemistry problem, a student has to evaluate the expression

$$\frac{4.00 \times 0.011 \times 273.16}{0.166 \times 299}$$

What figure should she quote as her answer?

Set 3: Random and systematic errors

Notes

All measurements have a degree of uncertainty resulting in experimental error. The experimental error in a result is the difference between the experimental value and the literature or theoretical value. There are two types of experimental error: random error and systematic error. Both should be considered when evaluating any quantitative investigation.

Random errors come from measurements that have an equal chance of being above or below the actual value.

Systematic errors are a result of flaws in the experimental method or apparatus that lead to a result that is always either above or below the true value.

Errors are discussed in detail on page ix and x.

Set 3: Exercises

- To perform a titration a student will require the use of a balance, a pipette, a burette and a volumetric flask.

List each of the following under the correct heading

Random error	Systematic error

- (a) incorrect etching of “the mark” on the volumetric flask
- (b) careless reading of the liquid levels in the burette
- (c) a fault in the balance such that it read 100 mg too heavy
- (d) allowing the pipette to drain for only 5 seconds instead of the recommended draining time of 30 seconds, following delivery of a volume of solution

- Four students were asked to read this measuring cylinder volume. John said 4.3 mL, Mary said 2.85 mL, Barry 4.15 mL and Kathy said 3.75 mL. Are any of them correct?



Set 3: Random and systematic errors

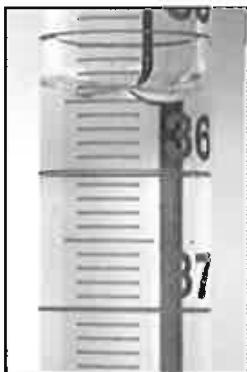
Notes

3. A student records a temperature as $17.00\text{ }^{\circ}\text{C}$ involving an error of $0.12\text{ }^{\circ}\text{C}$. He assumes that $0\text{ }^{\circ}\text{C}$ is exactly 273 K thereby making another error of 0.16 K . He uses his temperature value in a gas equation calculation. What is his maximum percentage error due to temperature?
4. A rectangular piece of domestic aluminium foil measuring $51\text{ cm} \times 3\text{ cm}$ has a mass of 3.0 g . What is the thickness of the foil if the density of aluminium is 2.70 g cm^{-3} ? ($\rho = m/v$)
5. Lyndon and Jenny were working as a team to complete a titration experiment. They obtained the following titre results

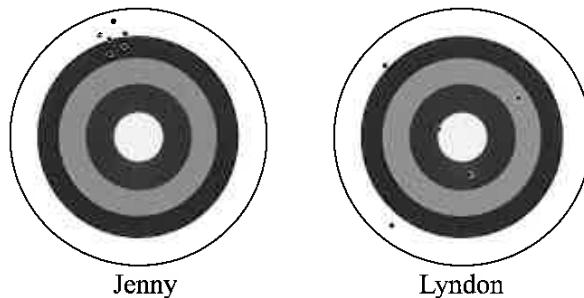
Final Reading (mL)	14.45	16.82	12.43	21.56	20.05
Initial Reading (mL)	34.57	37.34	32.98	42.01	40.59
Titre	20.12	20.52	20.55	20.45	20.54

Their teacher said the correct titre for the experiment was 22.54 mL .

- (a) Lyndon suggests they have a random error while Jenny suggests a systematic error. Who do you think is right?
(b) Jenny suggests repeating the experiment and averaging all the results. Do you agree?
(c) They cannot agree so they decide to abandon the experiment and start it all again. What changes would you suggest they make to achieve a successful result?
6. If the titration of 20.0 mL of 0.105 M NaOH with 0.098 M HCl yields an average titre of 20.9 mL calculate:
 - (a) the actual error
 - (b) the percentage error.
7. If a student observed a burette in this manner (left) would he be able to accurately take a reading? What is this mistake called? How do you know the reading will be inaccurate?
8. Two people are practicing archery. Their targets are shown below.



Question 7



- (a) Explain the terms 'random error' and 'systematic error' using the targets above.
- (b) Using the terms 'accurate' and 'precise' describe the results.
- (c) How could each person improve their result?

Use the following uncertainties for glassware information to complete questions 9 to 13:

Glassware	B grade tolerance	A grade tolerance
Burette (50 mL)	50 ± 0.10 mL	50 ± 0.05 mL
Pipette (20 mL)	20 ± 0.060 mL	20 ± 0.030 mL
Pipette (25 mL)	25 ± 0.060 mL	25 ± 0.030 mL
Volumetric flask (100 mL)	100 ± 0.20 mL	100 ± 0.10 mL
Volumetric flask (250 mL)	250 ± 0.30 mL	250 ± 0.15 mL
Volumetric flask (500 mL)	500 ± 0.50 mL	500 ± 0.25 mL

9. An analysis method consistently produces an absolute error of 0.6 mg. Calculate the percentage relative error caused by this uncertainty for the following sample masses.
 - (a) 600 mg
 - (b) 250 mg
 - (c) 30 mg
10. When titrating it is found that an over-titration of 0.05 mL is required to produce a visible colour change. Calculate the relative percentage error of this over titration if the titre is:
 - (a) 10.5 mL
 - (b) 25.3 mL
 - (c) 37.2 mL
11. 2.445 g of anhydrous sodium hydrogencarbonate is dissolved in distilled water, transferred to a 250.0 mL B grade volumetric flask and water is added to the mark. Calculate the concentration of the standard sodium hydrogencarbonate solution and the percentage and absolute uncertainties. (Assume the mass is accurate)
12. All glassware used was A grade. 20.00 mL aliquots of a 0.0446 mol L⁻¹ sodium hydrogencarbonate solution were titrated against a hydrochloric acid solution and the following titres obtained (in mL) 23.15, 22.45, 22.55, 22.60:
 - (a) Calculate the average titre and the percentage uncertainty based on the tolerance of the glassware from these results.
 - (b) Calculate the percentage uncertainty based on the range of the titres.
 - (c) Calculate the number of moles of sodium hydrogencarbonate and the percentage uncertainty in the conical flask (assume the initial concentration of sodium hydrogencarbonate is accurately known).
 - (d) Calculate the concentration of the hydrochloric acid solution and the percentage and absolute uncertainty.
13. Using B grade glassware, a sample of cloudy ammonia was titrated with 0.100 mol L⁻¹ hydrochloric acid. The titre values were found to be 5.44 mL, 5.60 mL, 6.00 mL and 5.55 mL.
 - (a) Calculate the average titre and the percentage uncertainty.
 - (b) In a second experiment, the cloudy ammonium was diluted by a factor of 5 and the following titres (in mL) obtained: 25.20, 24.50, 24.45 and 24.30.
 - (i) Calculate the average titre
 - (ii) Calculate the percentage uncertainty associated with the glassware.
 - (c) Using your answers to (a) and (b), explain why diluting the cloudy ammonia produces a more accurate result.

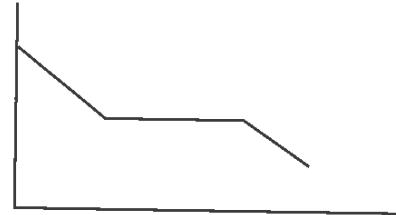
Set 4: Mixtures

Physical separation techniques are used to purify components or to determine the composition of a mixture. Separation techniques rely on the individual substances having different physical properties. These properties include solubility, level of attraction to a stationary phase, particle size, density, melting and boiling point.

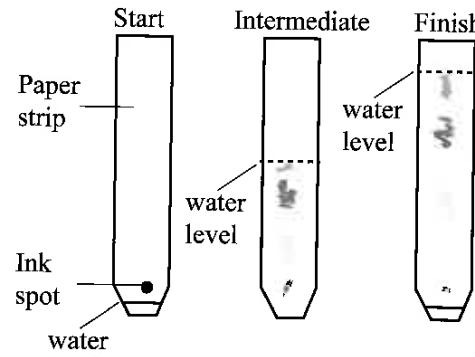
- Filtration relies on different particle size
- Distillation relies on different boiling points.
- Chromatography relies on the solubility of substances in a particular solvent and their levels of attraction to a stationary phase.

Set 4: Exercises

1. Describe using examples the differences between a homogeneous and a heterogeneous mixture.
2. The graph (right) shows the cooling curve of ethanol.
 - (a) Place the following labels on the graph: time, temperature, and the time during which freezing occurs.
 - (b) Research the freezing temperature of “pure” ethanol and add this to the graph. Describe the meaning of pure using ethanol as an example.
 - (c) Ethanol produced by distillation is not pure. State how this mixture should be described and name the other substance present.
3. The diagram (right) shows the chromatographic separation of the components of a sample of black ink.
 - (a) Describe how paper chromatography can be used to separate the different dyes in ink.
 - (b) Use the final chromatogram to describe the ink mixture.
 - (c) Which of the dyes show the least adhesion to the paper and the most solubility in water. Explain.
 - (d) Each separated dye can be assigned an R_f value. Describe what is meant by an R_f value and how it can be determined.
 - (e) State which colour dye has the largest R_f value.
4. Describe how you could separate and collect the first substance (in bold) from each mixture:
 - (a) **sucrose** and sand
 - (b) **sand** and sodium chloride
 - (c) **water** and copper sulphate
 - (d) **pink** and blue dyes in purple text ink.
 - (e) **octane** from crude oil
5. Explain the difference between simple distillation used to obtain water from salty water and fractional distillation to separate the components of crude oil.



Chromatographic separation of black ink



Notes

Set 5: Kinetic theory

Notes

Set 5: Exercises

1. In Experiment 7 you examined the relationship between the pressure and volume of a gas as described by Robert Boyle. State Boyle's Law, give a mathematical relationship between pressure and volume and sketch a graph to show this relationship.
2. Aerosol cans, including deodorant sprays, have labels warning against exposing used cans to excessive heat. Record the warning label of an aerosol can that you have at home. Explain the reasons behind this warning.
3. A car or bicycle tyre becomes hotter during use. Assuming the volume of the tyre remains constant how does the pressure of air inside the tyre change? Explain in terms of the Kinetic Theory of matter.
4. A helium balloon, when released, rises up into the air and disappears out of sight.
 - (a) Why does the helium balloon rise?
 - (b) Describe the pressure changes both in and around a helium balloon as it rises. Use diagrams to help your explanation.
 - (c) Why will the balloon eventually burst?
5. Using the Kinetic Theory explain why the boiling point of a liquid is a sharp definite temperature. Why does water boil at a temperature lower than 100 °C at 2 000 m above sea level?
6. Explain why liquids exert vapour pressure. Why is the boiling point of a solution higher than the boiling point of its solvent?

Set 6: Macroscopic properties of matter

Set 6: Exercises

Notes

1. Chemists have criteria for classifying matter as solid, liquid or gas, but some substances can fall into more than one category. Research the properties of substances such as glass, bitumen or supercritical carbon dioxide (used to make decaffeinated coffee). Write an argument for and against classifying each material into only one state of matter.
2. Electrolyte imbalances in the blood are closely examined during blood sample analysis. Abnormally high or low values of electrolytes such as sodium, potassium, magnesium or calcium ions can lead to vomiting and dehydration or more severe problems like kidney disease and heart failure. Draw up a table that summarises the different blood electrolytes, their importance and include their concentrations.
3. Explain the following observations:
 - (a) an aerosol container explodes when heated;
 - (b) a partly inflated balloon shrivels up when cooled in liquid nitrogen yet returns to its normal shape and volume when allowed to warm to room temperature.
4. Some cooks will tell you that adding a tablespoon of salt to a pot of water will make the boiling point of the liquid greater than 100 °C, and so your pasta is cooked faster than it would be in unsalted water. Investigate the validity of this statement, and suggest other methods that could be used to decrease the cooking time of pasta.
5. As a liquid is heated its vapour pressure above the liquid increases. When the vapour pressure reaches atmospheric pressure boiling occurs. Knowing this compare and contrast the boiling point of water in Perth (at sea level) to that at Kalgoorlie (about 500 m). What impact will this have on cooking in boiling water?
6. **Healthy water:**
 - (a) Why is dissolved oxygen in rivers and lakes important?
 - (b) How does the solubility of oxygen gas in water change when the water temperature is increased?
 - (c) What is thermal pollution?
 - (d) Discuss the impact of thermal pollution on the concentration of dissolved gases in natural waterways.
7. **SCUBA Diving:**

SCUBA diving, a popular underwater pastime, depends on a supply of compressed air to allow breathing under water. Compressed air is air kept under a pressure greater than atmospheric pressure.

 - (a) Why is the air in SCUBA tanks kept under pressure?
 - (b) List the gases added to the breathing tanks of a SCUBA diver and explain why they are used.
 - (c) Nitrogen is not included in the air in scuba tanks for deep and prolonged dives because at higher pressures a large quantity of nitrogen dissolves in the blood. The increased solubility of nitrogen in the blood at high pressures is responsible for the two conditions: *rapture of the deep* (nitrogen narcosis) and the *bends*. Explain these conditions.
8. Describe Charles' Law in terms of the kinetic theory.

Set 7: Elements and symbols

Many element names have ancient origins as they have a long history. Elements identified in the modern era are often named by the scientists who have isolated or created them for the first time. Any proposed names must be ratified by the International Union of Pure and Applied Chemistry (IUPAC) before they can appear on the Periodic Table.

Element symbols are a short hand form of the element name. All symbols consist of up to 3 letters, where the first letter is always a capital and the remaining letters are lowercase.

Some symbols represent the Latin name of an element rather than the name in current use. Sodium, for example, has the symbol Na, which is derived from sodium, its Latin name.

Notes

Your curriculum authority may provide a list of elements and common compounds that you must know.

Set 7: Exercises

1. Write the symbols of the following elements
 - (a) fluorine
 - (b) calcium
 - (c) manganese
 - (d) tungsten
 - (e) silver
 - (f) uranium
 - (g) platinum
 - (h) iodine
 - (i) neon
 - (j) ruthenium
 - (k) thorium
 - (l) astatine
 - (m) germanium
 - (n) technetium
 - (o) barium

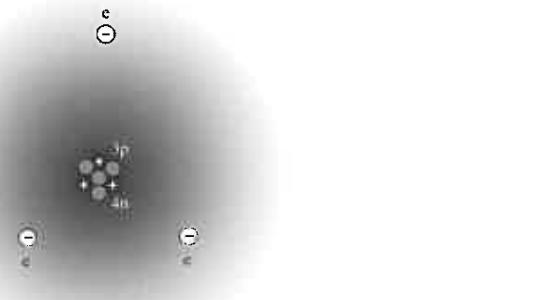
2. Write the names for the following symbols
 - (a) Na
 - (b) Re
 - (c) Cl
 - (d) Au
 - (e) Zr
 - (f) Pu
 - (g) Ce
 - (h) As
 - (i) Fe
 - (j) Co
 - (k) P
 - (l) Kr
 - (m) Zn
 - (n) K
 - (o) Sn

Set 8: Atoms and isotopes

Notes

Set 8: Exercises

1. a) Label the parts of the atom shown below. Note: This diagram is not to scale, the nucleus is hundreds of times smaller than shown.



- b) Complete the following using the atom shown.

Number of protons _____

Number of neutrons _____

Number of electrons _____

- c) Use a Periodic Table to identify the element. _____

2. Complete the table below:

Particle	Charge	Mass relative to a proton
proton		
neutron		
electron		

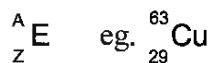
3. If an electron was removed from an atom:

- a) How would it affect the overall mass of the atom?
b) How would it affect the overall charge of the atom?

4. Copy and complete the sentence:

Isotopes of an element are atoms of the same element and hence have the same number of _____ and _____, but have different numbers of _____.

5. Isotopes are regularly represented by the following symbol :



For the element copper, identify:

E (element symbol): _____

A (mass number): _____

Z (atomic number): _____

6. Copy and complete the following table:

Symbol	Element	Mass No (A)	Atomic No (Z)	No of neutrons
$^{14}_6 \text{C}$		14	6	
	chlorine		17	18
$^{56}_{26} \text{Fe}$		56	26	
		31	15	
	silver	108		61
	carbon	12		
	sodium	23	11	
$^{64}_{29} \text{Cu}$		64	29	
	calcium	40		20
	carbon			7

7. Which element in the table in question 6 is represented by more than one isotope?
8. Use the format: $^{A}_{Z} \text{E}$ to represent the following isotopes of hydrogen.
- hydrogen-1
 - hydrogen-2
 - hydrogen-3

Set 9: Atomic structure and the Periodic Table

Notes

An electron configuration describes the arrangement of electrons surrounding the nucleus of an atom. Electrons occupy regions of space around the nucleus which we refer to as shells. Electrons in different shells have different energies. The energy of the electrons depends upon their distance from the nucleus and the shell they occupy. The electron configuration indicates how many electrons are in each energy level or shell.

Electrons, in their ground state, occupy the lowest energy levels or shells possible.

Examples

- Oxygen: atomic number (Z) = 8, there will also be eight electrons in a neutral atom of oxygen, two in the first energy level and the remaining six in the second energy level. The electron configuration of oxygen would be written as: 2, 6
- Oxide ion: O^{2-} , $Z = 8$, there will be eight electrons plus two extra electrons as it has a negative two charge. Therefore there will be two electrons in the first energy level and the remaining eight will be in the second energy level. The electron configuration of the oxide ion would be written as: 2, 8

Set 9: Exercises

- Copy and complete a table similar to the following to show details of the first 20 elements. As an example the element nitrogen has been done for you.

Z	Name	Symbol	Metal/non-metal	Electron configuration	Valence electron behaviour
1					
2					
3					
4					
5					
6					
7	Nitrogen	N	Non-metal	2, 5	Shares e^- or gains 3 e^-
8					

- (a) Which elements appear to always lose electrons?
(b) Describe the position of these elements on the Periodic Table.
- (a) Which elements appear to always gain electrons?
(b) Describe the position of these elements on the Periodic Table.
- (a) Which elements do not gain or lose electrons?
(b) Describe the position of these elements on the Periodic Table.
- Is there a pattern between the number of valence electrons and whether or not electrons are gained or lost?

6. Complete the following table:

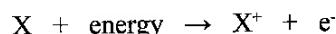
	Symbol	Atomic number	Number of protons	Number of electrons	Number of neutrons
Example	$^{19}_9\text{F}$	9	9	9	10
A	$^{16}_8\text{X}$				
B	$^{17}_8\text{X}$				
C	$^{16}_8\text{X}^{2-}$				
D	$^{17}_9\text{X}$				

7. In the table in question 6, which species are isotopes of each other? Explain your answer.
8. There is a saying that protons give an atom its identity and electrons give an atom its personality. Explain what this means.
9. Write the electron configurations for the following species:
 - (a) S²⁻
 - (b) Al³⁺
 - (c) K
 - (d) C
10. Which elements are represented by the following electron configurations?
 - (a) 2, 8, 7
 - (b) 2, 8, 8, 2
 - (c) 2, 8, 3
11. Which of the following electron configurations represent elements in which electrons are not in the ground state?
 - (a) 2, 7, 8, 1
 - (b) 1, 8, 8
 - (c) 2, 8, 7, 1

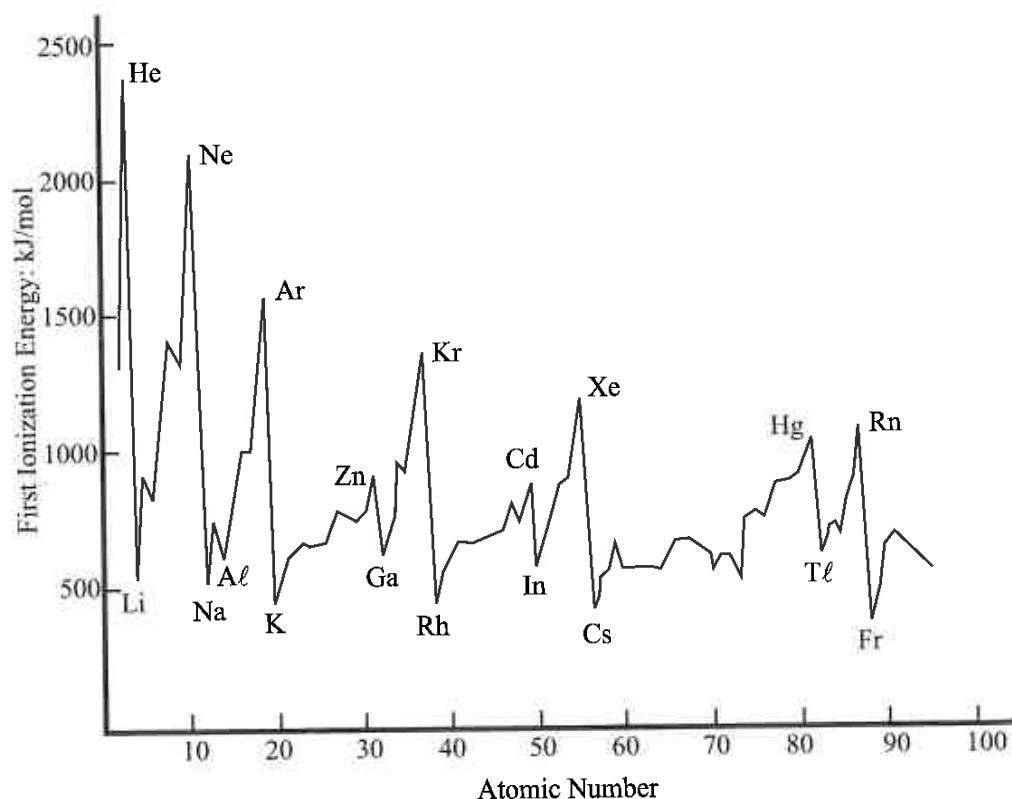
Set 10: Ionisation energy

Notes

Ionisation energy is the amount of energy required to remove one mole of electrons from one mole of atoms in the gaseous state. The process can be represented by:



The first ionisation energy is the energy required to remove the outer most electron from an atom forming a $1+$ ion. The second ionisation energy is the energy required to remove the next electron, forming a $2+$ ion. If the outer most electron is very loosely bound as in Group 1 and 2 elements the amount of energy required will be lower than for elements where the outer most electrons are more strongly bound ie. Group 17 and 18. The graph shows the first ionisation energy for elements in the Periodic Table.

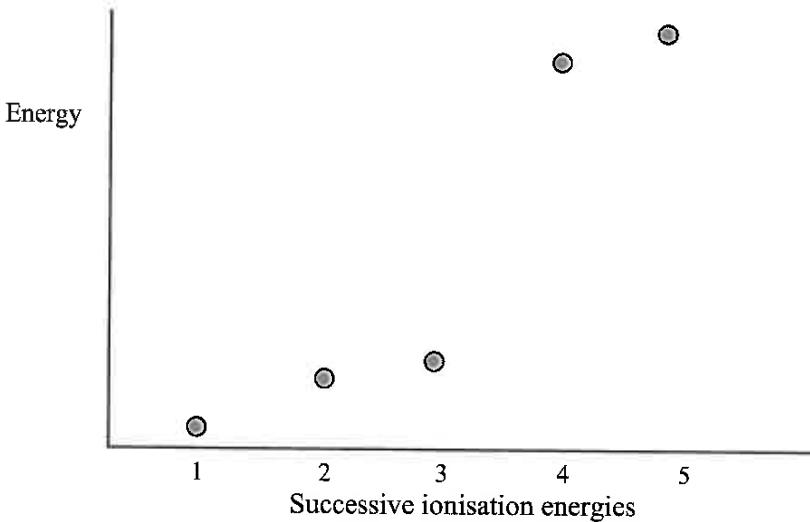


- The ionisation energies are highest for Group 18 elements and lowest for Group 1 elements.
- As you move across a Period the ionisation energies steadily increase.
- As you move down a Group the ionisation energies steadily decrease.

Set 10: Exercises

1. Explain what is meant by ionisation energy using a sodium atom to illustrate your answer.
2. Why would you expect an ionised atom to be in the gaseous state?
3. “The first ionisation energy of atoms shows periodicity” Explain what is meant by this statement.

4. Is there a relationship between ionisation energy and metal/non-metal properties of atoms? Explain your answer.
5. Explain why magnesium will not easily form 3+ ions.
6. Use the following graph showing consecutive ionisation energies for an unknown element to answer the questions.



- How many valence electrons does this element have?
- Justify your answer to (a).
- Will all the elements in the same Group as this element have similar shaped graphs? Explain.

Use the table below to answer questions 7, 8 and 9.

Element	1st ionisation energy (kJ mol ⁻¹)	2nd ionisation energy (kJ mol ⁻¹)	3rd ionisation energy (kJ mol ⁻¹)	4th ionisation energy (kJ mol ⁻¹)
A	418	3052	4420	5877
B	2080	3952	6122	9371
C	737	1450	7732	10542
D	899	1757	14849	21006
E	800	2427	3659	25025
F	590	1145	4912	6491

7. How many valence electrons does element F have? Explain your answer.
8. Which of these elements could be positioned before F in the Periodic Table? Explain your answer.
9. Which of these elements could be positioned after F in the Periodic Table? Explain your answer.
10. How can knowing the number of valence electrons help in determining the type of bonding that will occur in an element?

Set 11: Periodic trends

Notes

The Periodic Table that we use today was developed over many years and ordered all known elements into an arrangement that made it possible to make predictions about properties both physical and chemical.

Going from left to right the horizontal rows are called Periods. The vertical columns are called Groups.

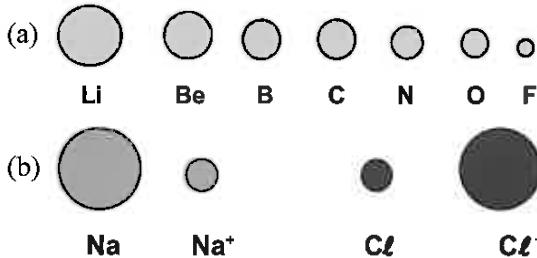
Periodic trends and patterns:

When going from left to right across a Period in the Periodic Table:

- Atomic radii decrease
- Electronegativity increases
- Number of valence electrons increases
- Ionisation energy increases
- Electrical and thermal conductivity decrease
- Elements change from metal to metalloid to non-metal
- Bonding in elements changes from metallic to covalent network to covalent molecular

Set 11: Exercises

1. List 5 general trends or patterns when you go down a Group on the Periodic Table.
2. Write two conclusions that can be made from the following diagram showing radii of atoms and ions.



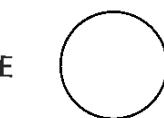
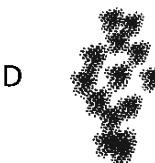
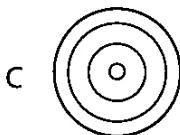
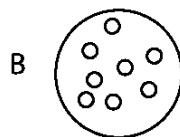
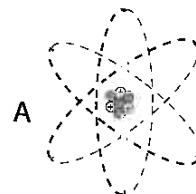
3. Explain the changes in the electronegativity of elements as you move:
 - (a) to the right across a Period of the table and
 - (b) up a Group on the table.
4. A new element is discovered and found to have 2 valence electrons. Write a paragraph describing the physical and chemical properties of this new element. What Group of the Periodic Table would you place it in?
5. The type of bonding changes as you move across the Periodic Table from metallic to covalent network to covalent molecular. Relate the changes in bonding to the number of valence electrons present.

Set 12: Properties and structures of atoms

Set 12: Exercises

1. The Atom Timeline: A student searches for images of the structure of the atom and finds many different images, including those on the right. Answer the following questions about the structure of the atom.
 - (a) Why are there different representations of the atom?
 - (b) What do the different images convey about the understanding of the structure of the atom?
 - (c) Why do you think the structure shown on page 126 is more accurate than structure A?
 - (d) Sequence the structures to make a timeline of our understanding of atomic structure. Include the structure on page 126.
 - (e) Briefly explain what evidence was used by scientists to determine the different structural representations shown.
2. Find out how emission spectroscopy is used to help identify plastics for recycling.
3. Describe how the following isotopes are used in medical research:
 - (a) Technetium-99
 - (b) Molybdenum-99
 - (c) Cobalt-60.
4. Each of the following scientists have made significant contributions to our understandings of the atomic model. For each scientist, list the year of their discovery and describe their contribution, including a diagram where appropriate.

Dalton
Thomson
Rutherford
Bohr
Chadwick
5. Under what circumstances would atomic absorption spectroscopy (AAS) be used?
6. With the aid of a diagram explain how an AAS works.



Set 13: Compounds and formulae

Notes

Chemical symbols and formulae are a shorthand way of communicating information about the structure and composition of elements and compounds. For a molecule the formula indicates what elements and how many of each type are present. So a molecule of carbon dioxide which is represented by the formula CO_2 consists of one carbon atom and two oxygen atoms. For metals, covalent network structures and ionic compounds the formulae indicate the ratio of elements present. So calcium chloride which is represented by the formula CaCl_2 consists of a lattice of positive calcium ions and negative chloride ions. For every calcium ion there are two chloride ions present, but the total number of ions is not known.

Writing the formulae of molecular compounds

Common names are often used for molecular compounds. Some common names and their formulae that you should know are shown in the table (below).

Common Name	Formula
Water	H_2O
Hydrogen peroxide	H_2O_2
Ammonia	NH_3
Methane	CH_4
Hydrogen chloride	HCl
Sulfuric acid	H_2SO_4
Phosphoric acid	H_3PO_4
Nitric acid	HNO_3
Acetic acid	CH_3COOH

For molecular compounds without common names the name of the compound usually includes information about the number of each element present. For example sulfur dioxide is written as SO_2 . Prefixes, listed in the following table, are used to show the number of atoms of each element.

Prefix	Number of atoms
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Rules for naming molecular compounds:	Example: CO ₂
1. The name of the element closer to the bottom left of the Periodic Table is written first.	Carbon is further to the left than oxygen so is named first
2. The name of the second element is changed to end with the suffix <i>-ide</i>	Oxygen is the second element so will be written as <i>oxide</i>
3. Prefixes are then used to identify the number of atoms of each element in the compound. (If there is only one atom of the first named element no prefix is required.)	Carbon – no prefix needed Oxygen – 2 atoms so use the prefix di - <i>Carbon dioxide</i>

Writing the formulae of ionic compounds

Writing the formula for an ionic compound requires knowledge of the charge on the ions present in the compound. As the overall charge in a compound must be zero, the charges on the ions present must add up to zero. The name only contains the name of the ions. Write the positive ion first.

Examples**1. Writing the formula of sodium fluoride**

Ions involved: Na⁺ F⁻

Use subscript numbers ~~Na₁⁺~~ ~~F₁⁻~~
to balance charge

Charges on the ions +1 -1 = 0
add to zero

**2. Writing the formula of calcium fluoride**

Ions involved: Ca²⁺ F⁻

Use subscript numbers ~~Ca₁²⁺~~ ~~F₂⁻~~
to balance charge

Charges on the ions +2 -2 = 0
add to zero



Set 13: Compounds and formulae

Notes

Set 13: Exercises

1. Name the following molecular substances

- (a) CO
- (b) SO₂
- (c) PCl₅
- (d) N₂S
- (e) P₂Br₄
- (f) SF₆

2. Write the formula of each of the following molecular compounds.

- (a) nitrogen monoxide
- (b) nitrogen dioxide
- (c) dinitrogen tetroxide
- (d) sulfur trioxide
- (e) water
- (f) pentaphosphorous decoxide
- (g) hydrogen chloride
- (h) hydrogen iodide
- (i) phosphorous tribromide
- (j) ammonia

3. Write the formula for each of the following ionic compounds.

- (a) lithium chloride
- (b) silver iodide
- (c) potassium nitrate
- (d) caesium ethanoate
- (e) barium bromide
- (f) copper(II) sulfate
- (g) manganese(IV) oxide
- (h) nickel nitrate
- (i) aluminium oxide
- (j) chromium(III) sulfate
- (k) lead(IV) phosphate
- (l) ammonium dichromate

4. Name the following compounds

- (a) CO
- (b) N₂O
- (c) H₂S
- (d) H₂O₂
- (e) H₃PO₄
- (f) Ca₃N₂
- (g) HNO₃
- (h) CoHPO₄
- (i) CuCl₂
- (j) FeSO₄

Set 14: Bonding and properties

Using the general properties of pure substances such as conductivity as solid, conductivity as liquid, melting point, hardness and brittleness, it is possible to see that most of these substances fall into one of four categories. We can link these properties to the types of bonding present within the substances. The four categories of substances are metallic, ionic, covalent network and covalent molecular. Each of these categories has a characteristic set of physical properties.

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Set 14: Exercises

1. From the physical properties listed in the table, predict the bonding class: metallic, ionic, covalent network or covalent molecular.

	Colour	Hardness	Electrical conductivity as solid	Electrical conductivity as liquid	Melting point (°C)
(a)	yellow	brittle	no	no	113
(b)	yellow	soft	yes	yes	1064
(c)	yellow	brittle	no	yes	402
(d)	white	soft	no	no	37
(e)	white	brittle	no	yes	801
(f)	white	brittle	no	no	146
(g)	brown	hard	yes	yes	1085

Set 15: Uses, properties and structure

Notes

By making observations about materials we may develop an awareness of the different properties of materials. The properties observed include colour, state, hardness, elasticity, conductivity, melting point and boiling point. Many of these properties can be explained by knowledge of the type of bonding that occurs in the material. Therefore, in order to understand the many materials that surround us it is essential to have an understanding of the structure of ionic, metallic, covalent network and covalent molecular substances. This will then enable us to develop an understanding of the relationships between the properties and structures of these substances.

The following sets of questions are designed to get you thinking about the relationship between structure and properties of materials.

Set 15: Exercises

1. Explain the following:
 - (a) electricians' screwdrivers have plastic handles
 - (b) saucepans have plastic or wooden handles
 - (c) plastic microwave dishes can't be used in a conventional oven
 - (d) tools used to etch glass have diamond tips
 - (e) lead is used to make sinkers for fishing
2. Lead is used in the making of leadlight windows. What properties of the metal make it suitable for this purpose?
3. Describe the property of gold that makes it suitable for the following uses:
 - a) jewellery
 - b) coinage
 - c) electronic circuits
 - d) shields in spacecraft
 - e) dental work
4. Classify the following as either metallic, ionic, covalent molecular or covalent network substances.
 - (a) water, H_2O
 - (b) copper(II) oxide, CuO
 - (c) methane, CH_4
 - (d) titanium dioxide, TiO_2
 - (e) iron, Fe
 - (f) silicon dioxide, SiO_2
5. Which of the substances listed in question 4 would conduct an electric current as a
 - (a) solid
 - (b) liquidExplain your choices.

Set 16: Properties and structures of materials

Set 16: Exercises

Notes

1. Research the composition of solder and compare the individual properties of the two components to the properties of the alloy.
2. How and why does the amount of carbon in steel change the properties of iron?
3. Why are carbon steel alloys more brittle and less malleable than the pure metal itself?
4. Perspex is said to remember it has been deformed for a while. Explain. Are there any metals or alloys with this property?
5. How can we use freezing point to decide if something is pure?
6. Nano-particles are very small – about 1/2000th the diameter of a human hair. Matter at the nanoscale can be manipulated to create new materials, composites and devices; the different characteristics of nanomaterials can be used to provide commercially available products. As products are designed on the basis of properties, which are different from the bulk material, their use can be associated with potential risks to health, safety and the environment and this has led to regulations being developed to address new and existing nanoform materials
 - (a) Identify and explain three potential benefits of using nano-particles.
 - (b) Some people have concerns about the safety of nano-particles. Identify and explain three problems that might be associated with their use.
7. Zinc oxide is commonly used as sunscreen in Australia. Zinc oxide sunscreen exists in both the nanoscale and non-nanoscale forms.
 - (a) For each form, list the advantages and disadvantages.
 - (b) What concerns have been raised about the use of zinc oxide sunscreens?
 - (c) What does the current research suggest about how safe it is to use nanoscale zinc oxide sunscreens?
8. Write an article suitable to be published in a weekend newspaper. Your article will discuss the relationship between the properties and structure of the different materials used to make outdoor furniture. In your article consider as many of the following as you feel necessary:
 - types of properties required of a material used to make outdoor furniture;
 - comparison of the physical and chemical properties of common materials such as wood, metal and plastic and discuss the suitability of each to make outdoor furniture;
 - innovations in chemical knowledge that has influenced the availability of wood, metal and plastic;
 - what you think outdoor furniture will be made of in the future;
 - cost;
 - availability.

Set 17: Electron dot diagrams

Notes

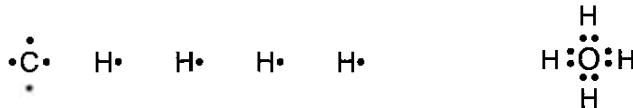
Electron dot diagrams are a useful tool to help you understand the bonding arrangements of atoms. From these diagrams it is possible to determine the shape of molecules and also their polarity. They are quite easy to draw for simple molecules as long as you follow some simple guidelines. Try to satisfy a full outer level of electrons (the octet rule) in all atoms except hydrogen (2 electrons) by:

- Determining the total number of valence electrons from all the atoms involved;
- Using a pair of electrons to form a single bond between each pair of bonded atoms; If there are not enough electrons, consider a multiple covalent bond.
- Placing the remaining electrons to satisfy the octet rule. Be sure the electrons are drawn in pairs.

Example 1: Methane CH₄

Carbon has four outer shell electrons and each of the hydrogen atoms has one electron, giving a total of 8 valence electrons.

- Trying a single covalent bond for each atom



A single bond between each carbon and hydrogen will satisfy all the elements.

Example 2: Carbon dioxide CO₂

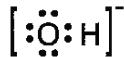
Each oxygen has six outer shell electrons and carbon has four outer shell electrons, giving a total of 16 valence electrons.

- Try a single covalent bond for each atom
Single bonds will not satisfy the carbon, there are not enough electrons.
- If there are not enough electrons, consider a multiple covalent bond. A double bond between each oxygen and carbon will satisfy all the elements.
- Place the remaining electrons to satisfy the octet rule. Be sure the electrons are drawn in pairs.



Example 3: Hydroxide ion OH⁻

Oxygen has six outer shell electrons and hydrogen has one outer shell electron. As a single negative ion, an extra electron has been gained, giving a total of 8 valence electrons.



A single bond between each oxygen and hydrogen will satisfy the elements, with the extra electron completing an octet for oxygen and is shown as a -1 charge on the ion.

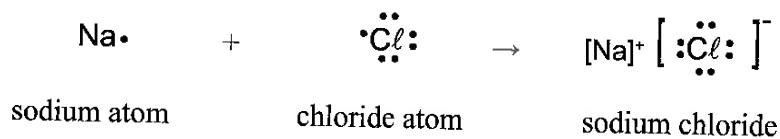
Example 4: Ionic substances

Notes

Square brackets are used when representing ions in electron dot diagrams of ions and ionic substances.

E.g. in NaCl , sodium has one valence electron and when forming an ionic bond becomes a positive ion by losing its valence electron.

Chlorine has seven valence electrons and accepts an electron when forming ionic bonds becoming a negative ion.



Set 17: Exercises

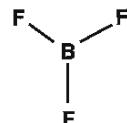
1. Draw electron dot diagrams for the following atoms and ions:
(a) Na
(b) Br
(c) P
(d) S^{2-}
(e) H^+
(f) N^{3-}
2. Draw electron dot diagrams for the following covalent molecular compounds (The central atom is in **bold** type) :
(a) H_2O
(b) CH_3Cl
(c) PH_3
(d) H_2S
(e) CHCl_3
(f) HCN
3. Draw electron dot diagrams for the following ionic compounds:
(a) NaOH
(b) CaCl_2
(c) Fe_2O_3
(d) K_2S
(e) Mg(OH)_2
(f) AgNO_3

Set 18: Molecular shape

Notes

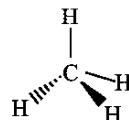
Polarity of a molecule is determined by the electron distribution in the bonds of the molecule and by the shape of a molecule. Shape is determined by the Valence Shell Electron Pair Repulsion Theory (VSEPR Theory). This states that electrons repel each other and will position themselves as far from each other as possible. The shapes you should be familiar with are shown in the diagram below.

Trigonal Planar



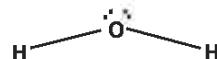
Bond angles for BF_3 are 120°

Tetrahedral



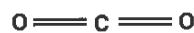
Bond angles for CH_4 are 109.5°

V-shaped



Bond angles for H_2O are 104.5°

Linear



Bond angles for CO_2 are 180°

Trigonal Pyramidal



Bond angles for NH_3 are 107°

There are two requirements for a molecule to be polar:

- The molecule should have one or more polar bonds and
- The molecule should not be symmetrical.

Bonds are polar if the elements bonded have differing electronegativities (electron attracting power).

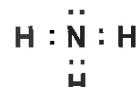
In the molecule of ammonia (NH_3) in the diagram above for example, nitrogen and hydrogen have differing electronegativities so polar bonds are present and the molecule shape is asymmetrical. This makes the ammonia molecule polar. Methane (CH_4) on the other hand has four polar bonds but the shape is symmetrical so the resulting molecule is non-polar.

Examples

When determining the shape and polarity of a molecule start by drawing an electron dot diagram.

1. Ammonia (NH_3): Nitrogen has 5 valence electrons. Hydrogen has one valence electron.

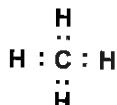
The electron dot diagram is:



Shape: As well as 3 bonding pairs of electrons there is a lone pair of electrons in the molecule and these also contribute to the shape of ammonia. The lone pair requires its share of space so the shape is **trigonal pyramidal** rather than trigonal planar.

Polarity: In the molecule of ammonia (NH_3), nitrogen and hydrogen have differing electronegativities so polar bonds are present. The molecule shape is trigonal pyramidal (asymmetrical). This makes the ammonia molecule **polar**.

2. **Methane (CH_4)**: Carbon has four valence electrons. Hydrogen has one valence electron. The carbon and hydrogen share electrons as shown in the electron dot diagram:



The molecule is symmetrical as the atoms space out evenly creating a **tetrahedral** shape. It has four polar bonds but the shape is symmetrical so the resulting methane molecule is **non-polar**.

Set 18: Exercises

- Explain why the electronegativity of iodine is much less than for fluorine.
- $\begin{array}{ccc} \text{Cl} & \cdots & \text{Cl} \\ \text{Cl} & \cdots & \text{H} \\ \text{Cl} & \cdots & \text{Na} \end{array}$
 - On the diagram above show by drawing a cross, where the shared pair of electrons would be most likely to be found between the two elements that are bonded. Explain why they would be in that position.
 - What is the significance of the position of the electrons to the bond formed?
- Which of the following species would have polar bonds?
 - I_2
 - CO_2
 - PH_3
 - CH_3Cl
 - O_2
- For the species in question 3, draw electron dot diagrams and determine the shape and the polarity of the molecules.
- Draw diagrams to represent each of the five shapes of molecules discussed in this set. Which shapes are always going to be polar? Explain why.
- Use examples to explain how it is possible to have a non-polar molecule with polar bonds.
- For the following species, draw the electron dot diagram and determine its shape.
 - SO_2
 - NH_3
 - H_2O
 - SO_3
 - NO_3^-
 - SO_4^{2-}
 - CO_3^{2-}
 - H_2O_2
 - C_2H_2
 - PCl_3

Set 19: Intermolecular forces

Notes

Observable properties of covalent molecular substances can be explained by the type and strength of intermolecular forces present between the molecules.

Dispersion forces are present between all molecules and are caused by temporary dipoles created by transitory imbalances in the electron cloud around a molecule. These temporary dipoles then induce dipoles in neighbouring molecules. The strength of dispersion forces are primarily determined by the number of electrons in the molecule, that is, the larger the number of electrons, the greater the strength of the dispersion forces. The surface area of the molecules will also impact on the strength of the dispersion forces. The greater the surface area, the greater the dispersion forces. This explains why isomers, with the same number electrons, have different boiling points.

Dipole-dipole forces of attraction will occur between polar molecules in addition to the dispersion forces. When comparing polar and non-polar molecules with similar numbers of electrons, differences in properties such as boiling point, can be explained by the presence of the additional dipole-dipole forces.

Hydrogen bonding is an extreme form of dipole-dipole forces. Where a very polar covalent bond between hydrogen and a very electronegative element such as fluorine, oxygen or nitrogen, exists in a molecule, the electron density around the hydrogen will be very low. As a result, the hydrogen atom will have a relatively high positive charge. The low electron density also allows the hydrogen atom to approach an unbonded pair of electrons on a neighbouring molecule's oxygen, nitrogen or fluorine (which are all relatively negatively charged) more closely. This combination of close proximity and relatively large difference in charge causes the dipole-dipole force of attraction to be much stronger than usual. This explains the higher than expected melting and boiling points in substances such as water, ammonia and hydrogen fluoride.

Set 19: Exercises

- For the following substances, identify the type(s) of intermolecular forces present between the molecules.

- (a) octane (C_8H_{18})
- (b) water (H_2O)
- (c) ethanol (C_2H_5OH)
- (d) pure nitric acid (HNO_3)
- (e) glucose ($C_6H_{12}O_6$) (see Figure 19.1)
- (f) iodine (I_2)
- (g) hydrogen chloride ($HC\ell$)
- (h) methane (CH_4)

- List the following substances in order of increasing boiling point and justify your choices.

- (a) HF, HI, $HC\ell$, HBr
- (b) H_2O , NH_3 , O_2 , CH_3Cl

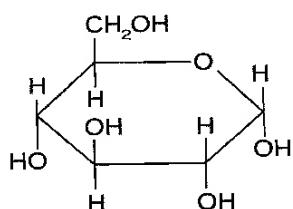


Figure 19.1: Glucose molecule

3. Given the following substances and their boiling points, list the substances in order of increasing vapour pressure at room temperature. Justify your choices.

Substance	Boiling point (°C)
H ₂ O	100
HNO ₃	83
Octane	125
Mercury	357

4. Explain how drawing an electron dot diagram can help you to predict the solubility of a substance.
5. Draw a dot diagram and predict the solubility of each of the following substances in water. Explain your decisions.
- (a) Methane (CH₄)
 - (b) ethanol (C₂H₅OH)
 - (c) ammonia (NH₃)
 - (d) hexane (C₆H₁₄)
 - (e) iodine (I₂)
 - (f) glucose (C₆H₁₂O₆) (see Figure 19.1)
6. Predict the solubility of each of the following substances in hexane (a non-polar liquid) and explain your decisions.
- (a) Methane (CH₄)
 - (b) ethanol (C₂H₅OH)
 - (c) ammonia (NH₃)
 - (d) hexane (C₆H₁₄)
 - (e) iodine (I₂)
 - (f) glucose (C₆H₁₂O₆) (see Figure 19.1)

Set 20: Chromatography

Notes

Set 20: Exercises

All forms of Chromatography are used to separate mixtures into their components.

In each case there is a stationary phase and a mobile phase. The components of the mixture show a degree of preference for either the mobile or stationary phase.

The stationary phase can be a solid (or a solid covered by a liquid) and the mobile phase can be a liquid or a gas.

Thin layer chromatography

1. As explained above, in all types of chromatography there is a mobile phase and a stationary phase. In Thin Layer Chromatography (TLC)
 - a) what is the mobile phase usually made of?
 - b) how is the stationary phase constructed?
2. A simple TLC experiment is to be set up using a plate and a mixture of dyes to be tested. The final result is shown In Figure 20.1.
 - a) Write a “method” for this experiment.
 - b) Why was a pencil rather than a ball point pen used to mark the starting position?
 - c) Why is the beaker covered with a lid?
 - d) Calculate the Rf factors for the red and the blue dye.
 - e) List two variables you controlled in this experiment

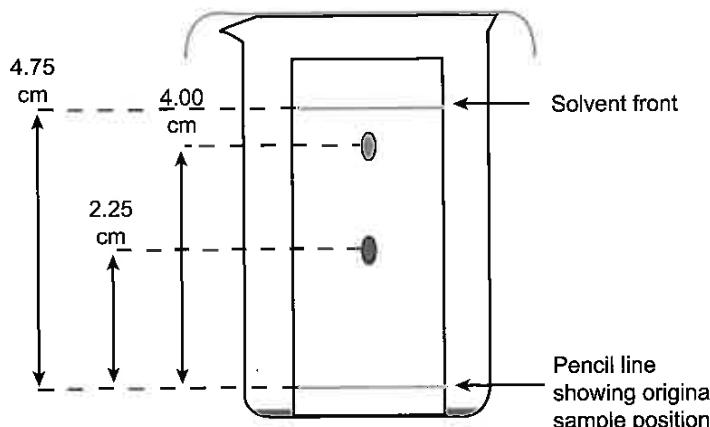


Figure 20.1: Glucose molecule

Gas chromatography (GC)

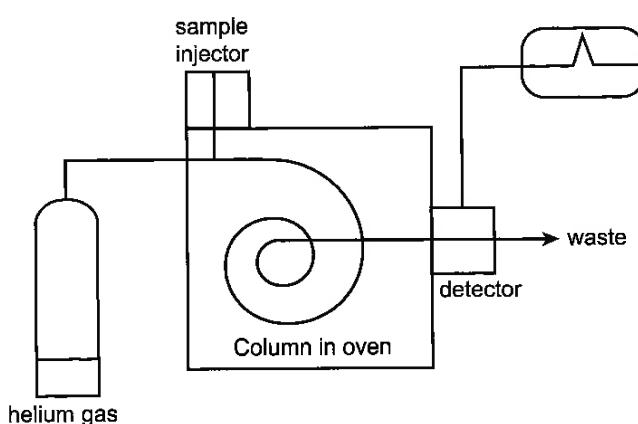
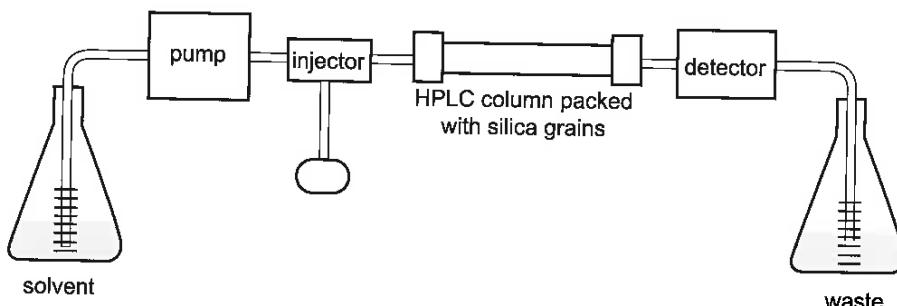


Figure 20.2: Gas chromatography

3. In gas chromatography (GC) the mobile phase is a gas (often helium) and the stationary phase is a liquid dispersed over a solid. The sample is either preferentially attracted to the gas or to the liquid on the solid material (column packing material). The time a sample takes to travel through the column is called its Retention Time – the more it is attracted to the liquid the longer its retention time.
 - a) Name the stationary phase and the mobile in gas chromatography.
 - b) Explain why a sample might have a very short retention time
 - c) If the temperature of the oven were raised why would the retention time decrease?

High-performance liquid chromatography (HPLC)

HPLC is a form of chromatography where, rather than allowing liquids to elute, or drip under gravity they are forced by a high pressure pump. Pressures of up to 40,000 kPa are often used. This means that the stationary phase can be very fine grained which gives an increased surface area and more interactions with the sample in the mobile phase. Hence it tends to produce better quality results.



4. As with GC the sample components are identified by their retention time. The retention time depends on the pump pressure, the size and chemical properties of the stationary phase, the solvent and its temperature.
 - a) Explain clearly the advantages of using a high pressure pump to force the solvent through the stationary phase.
 - b) Consider the diagram above of the HPLC system. How will the retention time for a polar molecule differ from a less polar molecule?
5. As in GC the retention time is calculated from when the sample is injected until it appears at the detector. List two factors that will determine the retention time.
6. Give three examples of mixtures that are commonly separated using high performance liquid chromatography (HPLC) and explain why this is the most appropriate method of separation for each mixture.
7. Give three examples of mixtures that are commonly separated using high performance liquid chromatography (HPLC) and explain why this is the most appropriate method of separation for each mixture.

Set 21: Relative atomic mass and mass spectroscopy

Relative atomic mass

Scientists have known for a long time that atoms of different elements have different masses. Originally the lightest element was given a value of 1 and all other elements were measured relative to that. As the mass was a relative value (a ratio), it had no units. Today, the masses of atoms have been more accurately determined and we know that an atom of hydrogen-1 (H-1) has a mass of 1.67353×10^{-27} kg. These tiny masses are awkward to work with and so scientists still often use relative masses.

The unified atomic mass unit (u), or dalton (Da), is defined as 1/12 the mass of an atom of C-12 in its ground state and has a mass of $1.6605402 \times 10^{-27}$ kg. Using this value, an atom of H-1 will have a mass of 1.0078 u. (<http://goldbook.iupac.org/U06554.html>)

Mass Spectrometry

Atomic mass spectrometry can be used to qualitatively and quantitatively identify elements present in a compound or mixture. These techniques can be used to measure concentrations as low as a few parts per billion. The basic principle that underlies the use of mass spectroscopy is that when charged particles move through a magnetic field they change direction. When the charge on the particles is the same, particles with the lower mass will experience a greater degree of deflection than heavier particles. When the mass is the same, the particles with a higher charge will experience a greater degree of deflection.

The first step in the process, vaporisation or atomisation, involves the separation of the substance into atoms in the gaseous phase. The atoms are then ionised, usually into +1 ions.

There are a number of ways of ionising a sample. The two methods most commonly used for mass spectroscopy are:

- Inductively coupled plasma (ICP)
- Electric spark

The ions produced will always be positively charged.

Electric sparks were originally used to produce the ions. These have been replaced by ICP. When ICP is used, the atoms are usually converted to singly charged ions.

After the sample has been ionised it is passed through a mass analyser. The stream of ions is accelerated through a magnetic field. The charged particles are deflected by the magnetic or electric field. The deflected particles then hit a detector which measures the number of particles and the amount of deflection. From this a mass spectrum is produced.

Mass spectra for chlorine showing the abundance of Cl-35 and Cl-37.

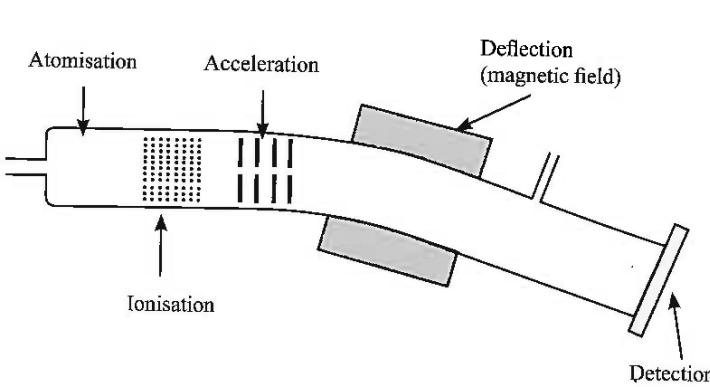
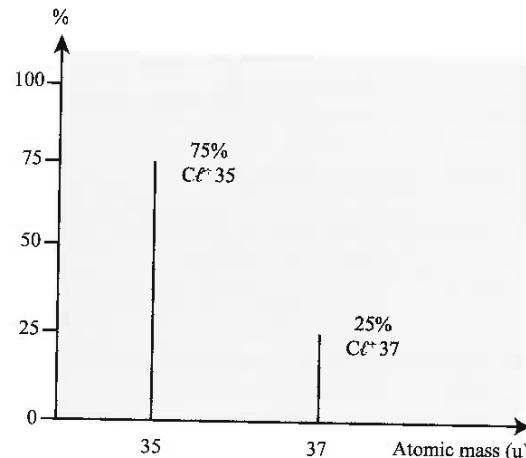


Figure 21.1: Mass spectrometer



Set 21: Relative atomic mass and mass spectroscopy

The separation is based on mass-to-charge ratio, so as long as the charge on the ions is +1, then the mass to charge ratio will be the same as the mass of the ion. For the exercise below, you may assume that the charge on the ions produced is always +1, although you need to be aware that ions with other charges can be produced.

The process of acceleration and deflection must occur in a vacuum as there should not be any gas particles that the ions could bounce off.

In summary, the process follows the steps atomisation, ionisation, acceleration, deflection and detection. The heavier the ion, the less it will be deflected within the magnetic field. While ions of different charge can be produced in a mass spectrometer, the most common charge is +1.

Calculating relative atomic mass from isotopic composition

Atoms of the same element with a different mass number or number of neutrons are called isotopes. The relative atomic mass and molar mass of elements are derived from the average weight of the various isotopes of an element. Knowing the relative isotopic masses and amounts of each isotope, the relative atomic mass can be calculated.

Atomic mass spectroscopy is an analytical technique that can be used to determine the relative abundances of the isotopes. The technique is described in detail elsewhere.

Examples

Boron consists of two main isotopes, Boron-11 and Boron-10. In a natural sample of boron, 80.1% will be Boron-11, while the remainder is Boron-10. Calculate the relative atomic mass of boron.

In order to calculate the relative atomic mass, each isotopic mass (which is close enough to the mass number for us to use that value) is multiplied by the percentage and divided by 100.

$$A_r(B) = \frac{10(19.9) + 11(80.1)}{100} = 10.8$$

Given the relative atomic mass and the isotopes, it is also possible to calculate the relative abundance of them. The simplest examples consist of two isotopes. To determine the relative abundance, simultaneous equations are used. The sum of the percentage abundances will always equal 100.

Boron consists of two isotopes, boron-10 and boron-11. The relative atomic mass of boron is 10.80. What is the percentage abundance of each isotope?

Let x be the percentage of boron-10 and y be the percentage of boron-11.

$$x + y = 100 \quad \text{equation 1}$$

$$10.80 = \frac{10(x) + 11(y)}{100} \quad \text{equation 2}$$

$$\text{Equation 1 can be rearranged: } y = 100 - x \quad \text{equation 3}$$

The value for y can then be substituted into equation 2

$$10.80 = \frac{10(x) + 11(100-x)}{100}$$

The equation can then be arranged to determine the value for x

$$1080 = 10x + 1100 - 11x \quad x = 20.0\%$$

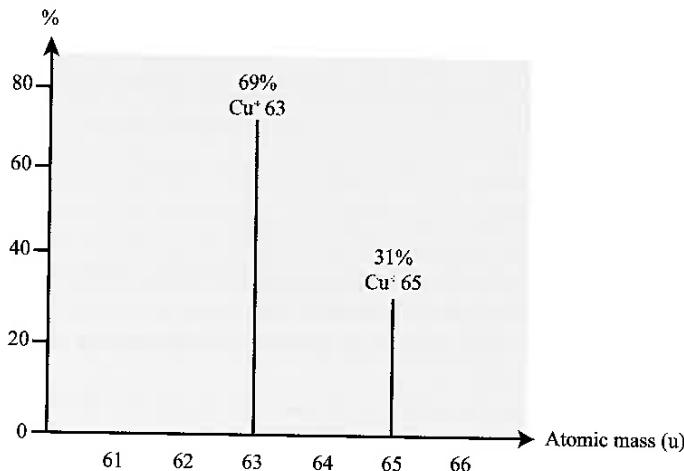
Substitute the value of x into equation 3

$$y = 100 - 20 = 80.0\%$$

Atomic mass spectroscopy spectra

A mass spectrum produced will plot the number of ions (relative amount) against the mass-to-charge ratio or mass where the ions are singly charged.

Notes



In the example above, a sample of copper was analysed. There are two isotopes of copper present. The peaks at 63 and 65 represent +1 ions. Based on this information, the relative atomic mass of copper can be calculated.

From the spectrum above, the percentage abundance of Cu-63 is 69 and the percentage abundance of Cu-65 is 31. Given this data, we can calculate the relative atomic mass of Cu:

$$\text{Ar}(\text{Cu}) = (69 \times 63) + (31 \times 65)/100 = 63.62$$

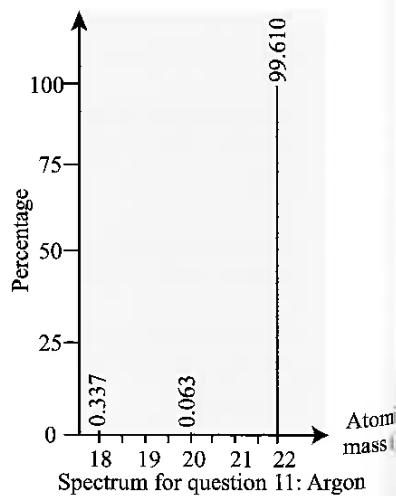
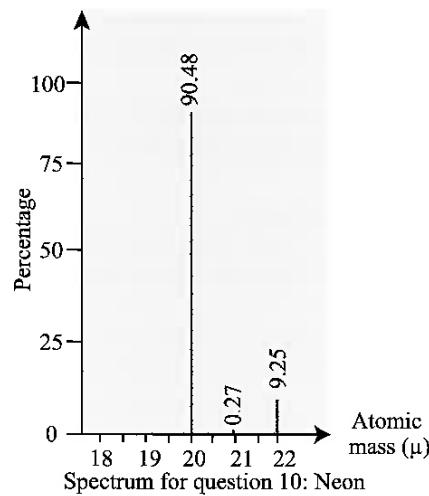
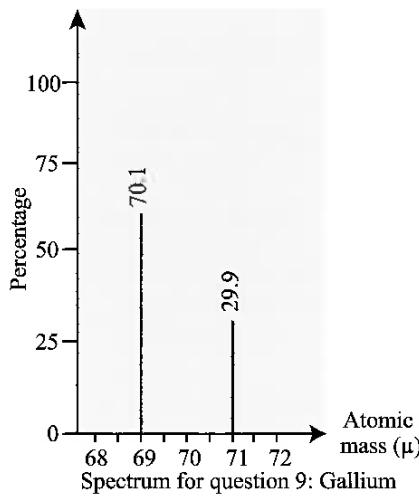
Set 21: Exercises

- Chlorine exists naturally as two isotopes, Cl-35, which accounts for 75.78% of all chlorine and Cl-37, which accounts for 24.22%. Calculate the relative atomic mass of chlorine.
- There three isotopes of silicon and their relative abundances are: Si-28 (92.2%), Si-29 (4.68%) and Si-30. Calculate the relative atomic mass of silicon.
- Four isotopes of lead and their relative abundances are Pb-204 (1.40%), Pb-206 (24.1%), Pb-207 (22.1%) and Pb-208 (52.4%). Calculate the relative atomic mass of lead.
- A particular atom of barium has a mass of 138 while the relative atomic mass of barium is 137.33. Explain the difference between these two values.
- Given the three isotopes of magnesium and their relative abundances of Mg-24 (79.0%), Mg-25(10.0%) and Mg-26, calculate the relative atomic mass of magnesium.
- Copper exists as two isotopes, Cu-63 and Cu-65. Given that the relative atomic mass of copper is 63.55, determine the percentage of each isotope present.

Set 21: Relative atomic mass and mass spectroscopy

Notes

7. Two samples of bromine, collected from different continents were found to have relative atomic masses of 79.9 and 80.9 respectively. What proportion of the stable isotopes of bromine, Br-79 and Br-81, are present in each sample?
8. Refer to the Figure 21.1 on page 149 to help answer the questions.
 - (a) Briefly describe each of the following processes that occur in a mass spectrometer
 - (i) ionisation
 - (ii) acceleration
 - (iii) deflection
 - (iv) detection
 - (b) The amount of deflection of a charged particle depends on the size of the charge and the mass of the particle.
 - (i) How does the size of the positive charge impact on the degree of deflection?
 - (ii) How does the mass of the particle impact on the degree of deflection?
 - (iii) Which of the following particles will experience the greatest deflection in each case?
I Cu⁺-63 or Cu²⁺-63
II B⁺-10 or B⁺-11
9. Gallium has two isotopes. The following mass spectra was produced. Calculate the relative atomic mass of Gallium. (Ga – 69: 70.1 %; Ga – 71: 29.9%)
10. The following mass spectra was produced from a sample of neon. Calculate the relative atomic mass of neon.
20 – 90.48% 21 – 0.27% 22 – 9.25%
11. The following mass spectra was produced from a sample of argon. Calculate the relative atomic mass of argon.
36 – 0.337% 38 – 0.063 % 40 – 99.610 %



Set 22: Molar mass

The mole concept

In order to perform calculations based on chemical reactions, we need to be able to express quantities of chemicals in terms of their mass. The mass of one atom, molecule or formula unit is too small to be practical; therefore a relative scale was devised. Masses of atoms, molecules or formula units are compared to the mass of an atom of carbon-12. This gives us ‘relative atomic mass’.

1 unified atomic mass unit = $\frac{1}{12}$ of the mass of 1 atom of $^{12}_6\text{C}$

In chemistry it is common usage to omit the unit when reporting relative atomic mass. Whilst this is useful, chemists found it more useful to determine the number of atoms in 12.00 g of carbon-12 and this number was designated one mole of atoms.

The mole concept is universally accepted and enables us to calculate molar masses of elements and compounds and perform stoichiometric calculations.

Notes

Molar mass

For ease of understanding we will refer to molar mass (M), which can be the mass of a mole of:

- atoms, eg Fe, Ne
- molecules, eg CO_2 , NH_3
- formula units, eg CuSO_4 , NaOH

The molar mass of a substance is numerically equal to the relative atomic, molecular or formula mass, but has the units g mol^{-1} .

The molar masses of elements are shown on the Periodic Table on the inside back cover.

Examples

Calculate the molar mass of the following giving your answer to 3 significant figures:

1. Magnesium (Mg): $M(\text{Mg}) = 24.31 \text{ g mol}^{-1}$

2. Ammonia (NH_3)

$$\begin{aligned} M(\text{NH}_3) &= M(\text{N}) + 3 \times M(\text{H}) \\ &= (14.01) + (3 \times 1.008) \\ &= 17.034 \\ &= 1.70 \times 10^1 \text{ g mol}^{-1} \end{aligned}$$

3. Alumina (Al_2O_3)

$$\begin{aligned} M(\text{Al}_2\text{O}_3) &= 2 \times M(\text{Al}) + 3 \times M(\text{O}) \\ &= (2 \times 26.98) + (3 \times 16.00) \\ &= 101.96 \\ &= 1.02 \times 10^2 \text{ g mol}^{-1} \end{aligned}$$

4. Copper(II) sulfate-5-water ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$)

$$\begin{aligned} M(\text{CuSO}_4 \cdot 5\text{H}_2\text{O}) &= M(\text{Cu}) + M(\text{S}) + 4 \times M(\text{O}) + 10 \times M(\text{H}) + 5 \times M(\text{O}) \\ &= (63.55) + (32.01) + (4 \times 16.00) + (10 \times 1.008) + (5 \times 16.00) \\ &= 249.64 \\ &= 2.50 \times 10^2 \text{ g mol}^{-1} \end{aligned}$$

Set 22: Molar mass

Notes

Isotopes

Most elements have several naturally occurring isotopes. This is why the molar masses of elements are not whole numbers. It is important that you are familiar with isotopes, both their structure and use.

Set 22: Exercises

1. Calculate the molar masses of
 - (a) potassium hydroxide, KOH
 - (b) copper(II) chloride, CuCl₂
 - (c) aluminium chloride, AlCl₃
 - (d) calcium hydroxide, Ca(OH)₂
 - (e) ammonium oxalate, (NH₄)₂C₂O₄
 - (f) sodium carbonate-10-water, Na₂CO₃·10H₂O
 - (g) tungsten carbide, WC
 - (h) methane, CH₄
 - (i) sucrose, C₁₂H₂₂O₁₁
 - (j) silver chloride
 - (k) zinc iodide
 - (l) sodium sulfate
 - (m) iron(III) sulfate
 - (n) sulfur dioxide
 - (o) sulfuric acid
2. Carbon dating is possible due to the composition of the C-14 isotope in living organisms.
Find out:
 - (a) how a C-14 atom is similar and different from a C-12 atom;
 - (b) any other isotopes of carbon;
 - (c) which isotope of carbon is the most abundant and how you can verify this;
 - (d) how carbon dating works.
3. Hydrogen has three isotopes; hydrogen, deuterium and tritium.
 - (a) Describe the composition of each of these atoms.
 - (b) Where are deuterium and tritium used?

Set 23: Moles, particles and mass

Notes

Atoms, molecules and ions are sub-microscopic with incredibly small masses. Yet we measure their mass and handle atoms in the laboratory all the time. So how do we do it? Small particles are dealt with in large numbers. We don't measure sugar and sand out as single grains. We use sensible units like teaspoons of sugar and truckloads of sand. Atoms are dealt with in a similar way where the sensible unit is called the mole. A mole of anything contains 6.022×10^{23} particles.

The number of atoms in 12.00 g of carbon-12 is 6.022×10^{23} . This very, very large number, 6.022×10^{23} is called Avogadro's number. Avogadro's number of particles is called a mole. Therefore one mole of any substance will contain 6.022×10^{23} atoms, molecules or formula units. The mole is a unit of measure used in chemistry.

For example:

- A mole of iron (Fe) consists of 6.022×10^{23} atoms of iron.
- A mole of carbon dioxide molecules (CO_2) consists of 6.022×10^{23} molecules of carbon dioxide.
- A mole of copper(II) sulfate (CuSO_4) consists of 6.022×10^{23} formula units of copper(II) sulfate.

Moles & mass

The relationship between the number of moles (n) and mass of a substance (m), in grams, is:

$$n = \frac{m}{M}$$

where M= molar mass of a substance in g mol^{-1} .

Examples

1. How many moles of calcium atoms are there in 30.0 g of calcium?

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{30.0}{40.08} & M(\text{Ca}) &= 40.08 \text{ g mol}^{-1} \\ n &= 7.49 \times 10^{-1} \text{ mol of Ca atoms} \end{aligned}$$

2. How many moles of benzene (C_6H_6) molecules have a mass of 390.0 g?

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{390.0}{78.107} & M(\text{C}_6\text{H}_6) &= 78.107 \text{ mol}^{-1} \\ n &= 4.99 \text{ mol of C}_6\text{H}_6 \text{ molecules} \end{aligned}$$

3. Calculate the mass of 5.0 moles of barium sulfate (BaSO_4) formula units.

$$\begin{aligned} n &= \frac{m}{M} \\ m &= nM \\ &= (5.0)(233.36) & M(\text{BaSO}_4) &= 233.36 \text{ g mol}^{-1} \\ m &= 1.2 \times 10^3 \text{ g} \end{aligned}$$

Set 23: Moles, particles and mass

Notes

Set 23: Exercises

1. Calculate the number of moles contained in each of the following:
 - (a) 72.0 g of magnesium (Mg)
 - (b) 4.00×10^2 g of calcium carbonate ($CaCO_3$)
 - (c) 104 g of ethyne (C_2H_2)
2. Calculate the mass of each of the following:
 - (a) 4.75 moles of lithium atoms
 - (b) 0.25 moles of sodium hydroxide formula units
 - (c) 9.00 moles of carbon monoxide molecules
3. Calculate the number of moles of molecules or formula units contained in each of the following:
 - (a) 28.0 g of nitrogen
 - (b) 232 g of butane (C_4H_{10})
 - (c) 3.90 g of sodium peroxide (Na_2O_2)
4. Calculate the number of moles of hydrogen peroxide (H_2O_2) molecules in 119 g of hydrogen peroxide.
5. Calculate the number of moles of sucrose ($C_{12}H_{22}O_{11}$) molecules in a can of cool drink that contains 39.8 g of sucrose.
6. Calculate the number of atoms in the following:
 - (a) 1.00 mole of gold (Au)
 - (b) 0.25 moles of sodium (Na)
 - (c) 2.33 moles of chlorine gas (Cl_2)
 - (d) 0.0894 moles of phosphorus (P_4)
7. Calculate the total number of ions in each of the following:
 - (a) 1.11 moles of sodium chloride ($NaCl$)
 - (b) 2.98 moles of aluminium oxide (Al_2O_3)
 - (c) 1.11×10^{-5} moles of magnesium fluoride (MgF_2)
 - (d) 0.222 moles of calcium sulfate ($CaSO_4$)
8. Calculate the total number of atoms in each of the following:
 - (a) 0.189 moles of barium phosphate ($Ba_3(PO_4)_2$)
 - (b) 25.0 g of bromine (Br_2)
 - (c) 0.0678 g of nitric acid (HNO_3)
 - (d) 12.5 g of acetic acid (CH_3COOH)

Set 24: Interpretation of formulae

The formula of a substance indicates the relative number of atoms of each element in the substance. For example:

Notes

One molecule of glucose, $C_6H_{12}O_6$, contains 6 atoms of carbon, 12 atoms of hydrogen and 6 atoms of oxygen.

Also one mole of glucose molecules, $C_6H_{12}O_6$ contains 6 mol of carbon atoms, 12 mol of hydrogen atoms and 6 mol of oxygen atoms.

Examples

- How many moles of oxygen atoms are contained in 5.00 moles of calcium phosphate ($Ca_3(PO_4)_2$) formula units?

$$\begin{aligned}n(\text{oxygen atoms}) &= 8 \times n(Ca_3(PO_4)_2) \\&= (8)(5.00)\end{aligned}$$

$$n(\text{oxygen atoms}) = 40.0 \text{ mol}$$

- What mass of oxygen atoms is contained in 132 g of carbon dioxide (CO_2)?

- (a) Find the number of moles of CO_2 using $n = \frac{m}{M}$ $M(CO_2) = 44.01 \text{ g mol}^{-1}$

$$\begin{aligned}n(CO_2) &= \frac{m}{M} \\&= \frac{132}{44.01} \\n(CO_2) &= 3.00 \text{ mol}\end{aligned}$$

- (b) Find the number of moles of oxygen from the number of moles of CO_2

$$\begin{aligned}n(O) &= 2 \times n(CO_2) \\&= (2)(3.00) \\n(O) &= 6.00 \text{ mol}\end{aligned}$$

- (c) Find the mass of O using $n = \frac{m}{M}$ $M(O) = (16.0 \text{ g mol}^{-1})$

$$\begin{aligned}n(O) &= \frac{m}{M} \\m(O) &= nM \\&= (6.00)(16.0) \\m(O) &= 96.0 \text{ g}\end{aligned}$$

Set 24: Interpretation of formulae

Notes

Set 24: Exercises

1. Calculate the number of moles of:
 - (a) hydrogen atoms in 2.50 moles of calcium hydroxide ($\text{Ca}(\text{OH})_2$) formula units
 - (b) nitrogen atoms in 0.0500 moles of ammonium nitrate (NH_4NO_3) formula units
2. Calculate the number of moles of:
 - (a) phosphorus trichloride (PCl_3) molecules which contain 21.0 moles of chlorine atoms
 - (b) potassium permanganate (KMnO_4) formula units which contain 2.00 moles of oxygen atoms
3. Calculate the mass of:
 - (a) oxygen in 795 g of copper(II) oxide (CuO)
 - (b) potassium in 1.04 g of potassium sulfate (K_2SO_4)
 - (c) calcium in 38.4 g of calcium oxalate (CaC_2O_4)
4. What mass of:
 - (a) sulfur dioxide (SO_2) contains 193 g of sulfur?
 - (b) ammonium sulfate ($(\text{NH}_4)_2\text{SO}_4$) contains 0.0960 g of hydrogen?
 - (c) octane (C_8H_{18}) contains 36.0 g of carbon?
5. Given 2.50 moles of ammonium phosphate ($(\text{NH}_4)_3\text{PO}_4$) formula units, calculate the number of moles of:
 - (a) ammonium (NH_4^+) ions
 - (b) phosphate (PO_4^{3-}) ions
 - (c) nitrogen atoms
 - (d) hydrogen atoms
 - (e) phosphorus atoms
 - (f) oxygen atoms
6. Calculate the mass of iron in 6.40×10^2 g of iron(III) oxide.
7. What mass of ethanol ($\text{C}_2\text{H}_5\text{OH}$) contains 144 g of carbon?
8. What mass of oxygen is contained in 2.50 moles of oxalic acid ($\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$)?
9. How many moles of nitrogen atoms are contained in 1.80×10^4 g of urea ($\text{CO}(\text{NH}_2)_2$)?
10. Calculate the mass of soap, sodium stearate ($\text{NaC}_{17}\text{H}_{35}\text{COO}$), that contains 1.25 g of carbon.

Set 25: Percentage composition

Notes

We are very interested today in knowing the exact composition of many of the products which surround us. When you buy foods, there are consumer panels listing the percentage of the different ingredients present. Medications also have the amount of active ingredient listed. There is increasing consumer demand that producers should supply specific information about goods. Food and medical goods are analysed by chemists to determine the exact percentage compositions of the ingredients. Individual compounds can also be analysed to determine their exact composition.

The percentage composition of a chemical compound specifies the percentage by mass of each of the different elements present in the compound.

$$\% \text{ mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100\%$$

Examples

- Calculate the percentage composition of ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$ to 3 significant figures.

$$\begin{array}{rcl} \% \text{ N} & = & \frac{28.02}{132.144} \times 100 = 21.2\% & \text{M(N)} & = & 14.01 \text{ g mol}^{-1} \\ \% \text{ H} & = & \frac{8.064}{132.144} \times 100 = 6.10\% & \text{M(H)} & = & 1.008 \text{ g mol}^{-1} \\ \% \text{ S} & = & \frac{32.064}{132.1} \times 100 = 24.3\% & \text{M(S)} & = & 32.06 \text{ g mol}^{-1} \\ \% \text{ O} & = & \frac{64.0}{132.144} \times 100 = 48.4\% & \text{M}((\text{NH}_4)_2\text{SO}_4) & = & 132.144 \text{ g mol}^{-1} \\ & & & & & \underline{100\%} \end{array}$$

Note: The percentages should add up to 100. The difference in the calculated value is due to rounding.

- Calculate the percentage of water of crystallisation in iron(II) sulfate-7-water, $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ to 3 significant figures.

$$\begin{array}{rcl} \% \text{H}_2\text{O} & = & \frac{\text{mass}(\text{H}_2\text{O})}{\text{M}(\text{FeSO}_4 \cdot 7\text{H}_2\text{O})} \\ \% \text{H}_2\text{O} & = & \frac{7(18.016)}{278.22} \times 100 & \text{M}(\text{FeSO}_4 \cdot 7\text{H}_2\text{O}) & = & 278.022 \text{ g mol}^{-1} \\ \% \text{H}_2\text{O} & = & 45.4\% & \text{M}(\text{H}_2\text{O}) & = & 18.016 \text{ g mol}^{-1} \end{array}$$

- A sample of soft solder consisted of 4.77 g of tin and 9.54 g of lead. Calculate the percentage composition of the solder to 3 significant figures.

$$\begin{array}{rcl} \% \text{ Sn} & = & \frac{\text{m (Sn)}}{\text{m (solder)}} \times 100 \\ & = & \frac{4.77}{(4.77 + 9.54)} \times 100 \\ & = & 33.3\% \\ \% \text{ Pb} & = & 100 - 33.3 = 66.7\% \end{array}$$

Set 25: Percentage composition

Notes

Set 25: Exercises

1. Calculate the percentage by mass of each element in:
 - (a) sodium hydroxide
 - (b) acetic acid (CH_3COOH)
 - (c) copper(II) sulfate-5-water
 - (d) potassium phosphate
2. Calculate the percentage by mass of:
 - (a) chlorine in calcium chloride
 - (b) sulfur in chromium(III) sulfide
 - (c) oxygen in potassium permanganate
 - (d) nitrogen in ammonium nitrate
3. Calculate the percentage by mass of water in:
 - (a) sodium carbonate-10-water
 - (b) nickel(II) sulfate-6-water
 - (c) barium chloride-2-water
4. An alloy is prepared by melting together 25.44 g of bismuth, 15.36 g of lead and 7.20 g of tin.
 - (a) Calculate the percentage composition of the alloy.
 - (b) Calculate the mass of each metal required to make 150.0 g of the alloy.
5. An alloy used in aircraft construction consists of aluminium, copper and magnesium. An 11.34 g sample of the alloy was treated with alkali to dissolve the aluminium, leaving a residue of mass 2.73 g. This residue was treated with dilute hydrochloric acid solution to dissolve the magnesium. The remaining residue had a mass of 0.900 g. Calculate the percentage composition of the alloy.
6. A 3.030 g sample of zinc was heated in oxygen to produce zinc oxide. The mass of zinc oxide produced was 3.771 g. Calculate the percentage composition of the zinc oxide.
7. A 16.00 g sample of copper(II) oxide was reduced by reaction with hydrogen gas to yield 12.77 g of pure copper. Calculate the percentage composition of the original copper(II) oxide.
8. Copper is found in such minerals as chalcopyrite (CuFeS_2) and malachite ($\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$).
 - (a) Calculate the percentage by mass of copper in each.
 - (b) What mass of chalcopyrite must be smelted to produce 1.00×10^2 kg of copper?
9. Darling Range bauxite contains high levels of the mineral gibbsite ($\text{Al}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$) with kaolinite clay ($\text{Al}_2\text{O}_3 \cdot 2\text{SiO}_2 \cdot 2\text{H}_2\text{O}$). Calculate the percentage of aluminium and the percentage of water in each of these substances.
10. Ilmenite from Capel contains approximately 53.5% titanium dioxide (TiO_2).
 - (a) Calculate the percentage of titanium in titanium dioxide.
 - (b) What mass of ilmenite must be refined to produce 1.00 tonne of pure titanium metal?

Set 26: Gas volumes

From our understanding of Kinetic Theory we know that gas volumes are independent of the identity of the gas. Gas volume is dependent upon the number of particles, its pressure and volume. This means that, for an ideal gas, one mole of any gas will occupy the same volume under the same pressure and temperature conditions. This enables us to develop a relationship between volume and number of moles of a gas.

Notes

Molar volume of gases

The molar volume of a gas is the volume occupied by 1 mole of the gas. The molar volume for an ideal gas, and for some real gases, is 22.71 L at S.T.P. (100 kPa, 25 °C). The relationship between the number of moles (n) of a gas and its volume (V) in litres at S.T.P. is

$$n = \frac{V(\text{in litres at S.T.P.})}{22.71}$$

Examples

- Calculate the volume occupied by 2.75 moles of carbon monoxide at S.T.P.

$$n = \frac{V}{22.71}$$

$$\begin{aligned} V &= n \times 22.71 \\ &= (2.75)(22.71) \\ V &= 62.5 \text{ L} \end{aligned}$$

- What is the volume of 9.60 g of oxygen at S.T.P.?

(a) Calculate the number of moles.

$$\begin{aligned} n(O_2) &= \frac{9.60}{32.0} \quad M(O_2) = 32.0 \text{ g mol}^{-1} \\ &= 0.300 \text{ mol} \end{aligned}$$

(b) Find the volume at S.T.P.

$$\begin{aligned} n &= \frac{V}{22.71} \\ V &= n \times 22.71 \\ &= (0.300)(22.71) \\ V &= 6.81 \text{ L at S.T.P.} \end{aligned}$$

Set 26: Gas volumes

Notes

Set 26: Exercises

1. Calculate the volume occupied by:
 - (a) 6.50 moles of carbon dioxide at S.T.P.
 - (b) 0.850 moles of hydrogen at S.T.P.
2. Calculate the number of moles of:
 - (a) methane (CH_4) in 4.50 L of the gas at S.T.P.
 - (b) oxygen in 25.0 mL of the gas at S.T.P.
3. Calculate the volume of:
 - (a) 100 g of propane (C_3H_8) at S.T.P.
 - (b) carbon dioxide gas at S.T.P. produced when a 1.00 kg block of dry ice, which is solid CO_2 , sublimes.
4. Calculate the volume occupied by the following masses of gases at S.T.P.
 - (a) 1.34 g of methane (CH_4)
 - (b) 2.00 g of ethane (C_2H_6)
 - (c) 58.6 g of ammonia (NH_3)
 - (d) 0.566 g of sulfur dioxide (SO_2)
5. 4.18 g of a gas occupies 1.00 L at S.T.P. Calculate the molar mass of the gas.
6. 3.49 g of a gas occupies 1.00 L at S.T.P. Calculate the molar mass of the gas.
7. What mass of oxygen gas will occupy 1.00 L at S.T.P.?
8. What masses will the following volumes of gas weigh?
 - (a) 2.55 L of sulfur trioxide (SO_3)
 - (b) 89.5 L of nitrogen (N_2)
 - (c) 0.00253 L of argon
 - (d) 41.2 L of chlorine (Cl_2)

Set 27: Ionic equations

Notes

Ionic equations are written for reactions that take place in solution. They represent only the species that are involved in the reaction. The following rules apply to the writing of ionic equations.

1. Strong electrolytes are written in ionic form.
2. Weak electrolytes are written in molecular form.
3. Non-electrolytes are written in molecular form.
4. Insoluble substances are written as the formulae of the substances.
5. Gases are written in molecular form (except for noble gases).
6. Ionic equations should not include species that do not change (spectator ions).
7. Equations must be balanced in terms of atoms and electrical charge.

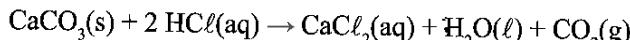
Knowledge of solubility rules is invaluable when writing ionic equations.
See the table solubility rules on page 207.

Examples

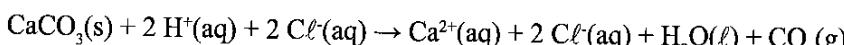
1. Write a balanced ionic equation for the dissolution of solid NaCl in water. NaCl is a solid, which, when dissolved, dissociates into ions as follows:



2. Write a balanced ionic equation showing what reaction occurs when dilute hydrochloric acid is added to solid calcium carbonate. Hydrochloric acid contains H⁺ and Cl⁻ ions since it is a strong electrolyte. The products of adding an acid to a metal carbonate are a salt, water and carbon dioxide.
(a) Write a balanced formula equation (including state symbols is useful):

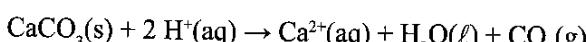


- (b) Use the rules for writing ionic equations to write relevant species as ions:



Note: Calcium chloride is a soluble salt.

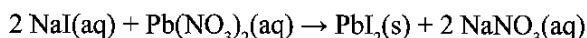
- (c) Omit spectator ions and write the final balanced ionic equation:



Set 27: Ionic equations

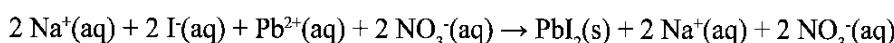
Notes

3. Write a balanced ionic equation showing the products formed when a solution of sodium iodide is added to a solution of lead(II) nitrate.
(a) Write a balanced formula equation:

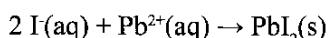


Note: Lead(II) iodide is an insoluble salt whereas sodium nitrate is a soluble salt.

- (b) Write relevant species as ions:



- (c) Omit spectator ions and write the balanced ionic equation:



Set 27: Exercises

Write balanced ionic equations for the following reactions:

1. The dissolving of potassium chloride.
2. The dissolving of barium nitrate.
3. The dissolving of sodium hydroxide.
4. Dilute nitric acid added to a solution of potassium hydroxide.
5. Zinc oxide powder added to sulfuric acid solution.
6. Dilute nitric acid being added to solid calcium carbonate.
7. A piece of clean lead added to a solution of silver nitrate.
8. The addition of a silver nitrate solution to a solution containing magnesium chloride.
9. Carbon dioxide gas being bubbled through limewater solution.
10. Phosphoric acid added to solid silver carbonate.
11. Hydrogen sulfide gas bubbled through silver nitrate solution.
12. Mixing of nitric acid and sodium carbonate solutions.
13. Sodium iodide solution added to lead(II) acetate solution.
14. The addition of excess sodium hydroxide solution to a suspension of aluminium hydroxide in water to produce a sodium tetrahydroxyaluminate ($\text{Na}[\text{Al(OH)}_4^-]$) solution.

Set 28: Equations and observations

Notes

An important skill in chemistry is the ability to predict the products when given some reactants. This requires knowledge of the general reaction types. From the products, an equation can be written and the observations expected can be described. Observations include the dissolving or formation of solids (with colours), evolution of gases (including colours and odour) and changes in solution colour. They may also include the formation or absorption of heat and even the production of light and sound.

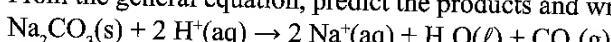
Example

Solid sodium carbonate is added to dilute hydrochloric acid solution.

Step 1 Identify the type of reaction.

This is an acid + carbonate reaction and has the general equation
Acid + carbonate → salt + water + carbon dioxide

Step 2 From the general equation, predict the products and write the ionic equation.



Step 3 Describe the observations expected

White solid dissolves; colourless, odourless gas evolved; colourless solution formed.

Set 28: Exercises

For the following reactants, write balanced equations and describe the expected observations.

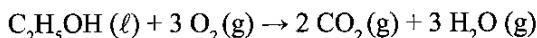
1. Marble chips (calcium carbonate) are added to dilute nitric acid.
2. Magnesium metal is added to dilute hydrochloric acid.
3. Baking soda (sodium hydrogencarbonate) is added to vinegar (acetic acid).
4. Sodium hydroxide solution is added to sulfuric acid solution.
5. Solid sodium hydroxide is added to dilute hydrochloric acid.
6. Dilute sulfuric acid is added to solid cobalt carbonate.
7. Barium chloride solution is added to dilute sulfuric acid.
8. Lead(II) nitrate solution is added to sodium iodide solution.
9. Gold(III) chloride solution is added to solid copper.
10. Sodium metal is added to water.
11. Potassium carbonate solid is added to an excess of dilute nitric acid.
12. Potassium hydroxide solution is added to nitric acid solution.
13. Silver nitrate solution is added to sodium bromide solution.
14. Nickel metal is added to copper nitrate solution.
15. Iron(II) sulfate solution is added to potassium hydroxide solution.

Set 29: Stoichiometry

Notes

A driver has been stopped by the police as part of a random breath test. There are a number of analytic techniques that can be used to determine the amount of alcohol in the driver's blood. Many of these techniques rely on reacting the alcohol with other chemicals and measuring the amount of products formed. In order to determine the amount of alcohol an understanding of the relationship between the reactants (including alcohol) and the products in the reaction is required. This relationship, expressed with the aid of a chemical equation, is called stoichiometry.

For example the equation representing the burning of ethanol is:



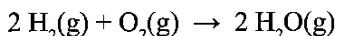
This indicates that one mole of $\text{C}_2\text{H}_5\text{OH}$ reacts with three moles of O_2 to form two moles of CO_2 and three moles of H_2O .

Using reaction stoichiometry, the relationship between the number of moles, masses and gaseous volumes of reactants and products can be calculated. We will start with the relationship between masses and moles.

Examples

- When a mixture of hydrogen and oxygen is sparked an explosive reaction occurs in which water is formed. If 0.480 moles of hydrogen gas is mixed with excess oxygen and sparked, calculate:
 - the number of moles of water formed
 - the mass of water formed
 - the number of moles of oxygen consumed
 - the mass of oxygen consumed.

- (a)(i) Write a balanced equation:



- (ii) Identify the unknown and relate it to the quantity which is known:

$$\begin{aligned} n(\text{H}_2\text{O}) &= n(\text{H}_2) \\ &= 0.480 \text{ mol} \end{aligned}$$

- (b) Convert the moles to mass ($m=nM$):

$$\begin{aligned} n(\text{H}_2\text{O}) &= 0.480 \\ m(\text{H}_2\text{O}) &= 0.480 (18.0) \quad M(\text{H}_2\text{O}) = 18.0 \text{ g mol}^{-1} \\ &= 8.60 \text{ g} \end{aligned}$$

- (c) Relate the unknown to the known (using the equation):

$$\begin{aligned} n(\text{O}_2) &= \frac{1}{2} \times n(\text{H}_2) \\ &= (\frac{1}{2})(0.480) \\ &= 0.240 \text{ mol} \end{aligned}$$

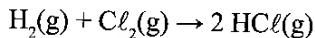
(d) Convert moles to mass ($m = nM$):

$$n(O_2) = 0.240$$

$$\begin{aligned} m(O_2) &= (0.240)(32.0) \quad M(O_2) = 32.0 \text{ g mol}^{-1} \\ &= 7.68 \text{ g} \end{aligned}$$

2. 17.8 g of chlorine gas reacts with hydrogen to form hydrogen chloride gas.
Calculate:
(a) the number of moles of hydrogen chloride formed and;
(b) the mass of hydrogen chloride formed.

(a)(i) Write a balanced equation:



(ii) Identify the unknown and relate it to the quantity which is known:

$$\begin{aligned} n(HCl) &= 2 \times n(Cl_2) \\ &= 2 \times \frac{17.8}{70.9} \quad M(Cl_2) = 70.9 \text{ g mol}^{-1} \end{aligned}$$

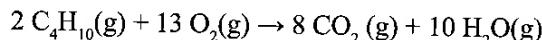
$$n(HCl) = 0.502 \text{ mol}$$

(b) Convert the moles to mass

$$\begin{aligned} n(HCl) &= 0.502 \\ m(HCl) &= (0.502)(36.458) \quad M(HCl) = 36.458 \text{ g mol}^{-1} \\ &= 18.3 \text{ g} \end{aligned}$$

Set 29: Exercises

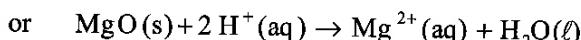
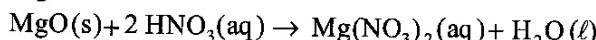
1. The equation for the combustion of butane is:



Calculate the number of moles of:

- (a) CO_2 produced in the combustion of 1.00 moles of C_4H_{10}
- (b) H_2O produced in the combustion of 3.00 moles of C_4H_{10}
- (c) O_2 consumed in the combustion of 0.600 moles of C_4H_{10}

2. 0.0300 moles of magnesium oxide is completely reacted with nitric acid to form a solution of magnesium nitrate:



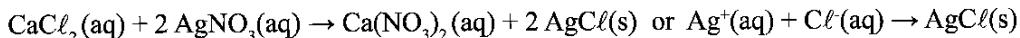
Calculate the number of moles and mass of:

- (a) nitric acid required
- (b) magnesium nitrate formed

Set 29: Stoichiometry

Notes

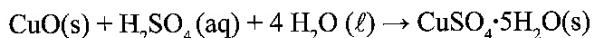
3. When silver nitrate solution is added to a solution of calcium chloride, a white precipitate of silver chloride is produced



If 0.200 moles of silver chloride is formed, calculate the number of moles and masses of :

- (a) silver nitrate required
- (b) calcium chloride required
- (c) calcium nitrate formed in solution

4. A sample of copper(II) oxide was reacted with sulfuric acid and the solution evaporated to dryness to yield 3.14 g of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$:

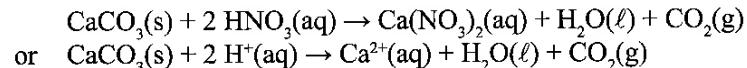


- (a) Calculate the mass of sulfuric acid required.
- (b) Calculate the moles of copper(II) oxide dissolved.

5. Write a balanced equation for the decomposition of potassium chlorate (KClO_3) into potassium chloride and oxygen gas by heating. How many moles of oxygen would be formed from the decomposition of 0.800 mol of KClO_3 ?

6. Write a balanced equation for the decomposition of mercury(II) oxide into mercury and oxygen gas (O_2). What mass of oxygen would be formed from the decomposition of 0.240 mol of mercury(II) oxide?

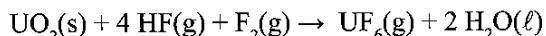
7. Excess nitric acid was added to 3.00 g of calcium carbonate:



Calculate the masses of:

- (a) nitric acid consumed
- (b) carbon dioxide produced
- (c) calcium nitrate formed in solution

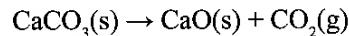
8. In the processing of uranium, one of the steps involves converting UO_2 to UF_6 .



For 7.50 kg of UO_2 , calculate:

- (a) the mass of hydrogen fluoride required
- (b) the mass of fluorine required
- (c) the mass of UF_6 produced

9. A load of quarried limestone contains 92.0% CaCO_3 , the remainder being silica (SiO_2). The limestone is to be heated to form quicklime (CaO).



- (a) What mass of quicklime could be produced by heating 5.00×10^2 kg of the limestone?
- (b) What mass of carbon dioxide would be produced?

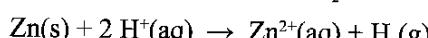
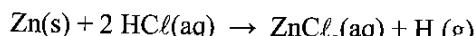
Set 30: Stoichiometry and gas volumes

Example

Determine the volume of hydrogen (measured at S.T.P.) that is produced when 45.6 g of zinc metal is added to excess hydrochloric acid.

Notes

- (i) Write an equation for the reaction



- (ii) Calculate the number of moles of zinc

$$n(\text{Zn}) = \frac{m}{M}$$

$$= \frac{45.6}{65.38}$$

$$= 0.697 \text{ mol}$$

- (iii) From the equation determine the number of moles of hydrogen

$$n(\text{H}_2) = n(\text{Zn})$$

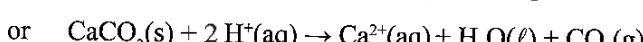
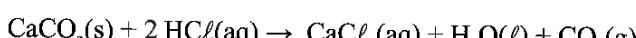
$$n(\text{H}_2) = 0.697 \text{ mol}$$

- (iv) Calculate the volume of hydrogen

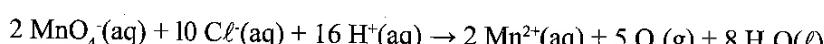
$$\begin{aligned} V &= n \times 22.71 \\ &= 0.697 \times 22.71 \\ &= 15.8 \text{ L} \end{aligned}$$

Set 30: Exercises

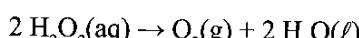
1. Calculate the volumes of gas at S.T.P. produced in each of the following reactions:
- (a) When 0.250 moles of CaCO_3 is treated with excess hydrochloric acid:



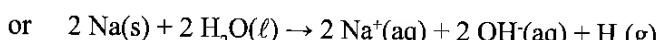
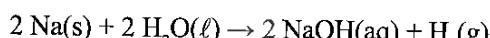
- (b) When 0.150 moles of potassium permanganate reacts with excess concentrated hydrochloric acid according to the equation



- (c) When 0.0300 moles of hydrogen peroxide disproportionates according to the equation



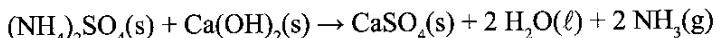
2. If 4.60 g of sodium reacts with water, calculate the volume of hydrogen produced at S.T.P.:



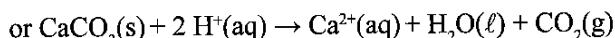
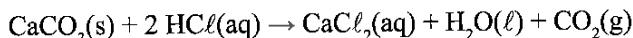
Set 30: Stoichiometry and gas volumes

Notes

3. Calculate the volume of ammonia produced at S.T.P. when 22.8 g of ammonium sulfate is heated with excess moist calcium hydroxide:

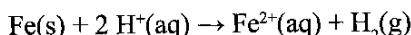


4. When 25.9 g of impure limestone was treated with excess hydrochloric acid, 5.6 L of carbon dioxide was produced at S.T.P.:



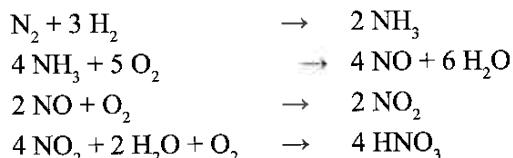
Calculate the percentage purity of the limestone.

5. A 15.3 g piece of steel, containing only iron and carbon, was treated with excess hot acid to form 6.03 L of hydrogen at S.T.P.:



Calculate the percentage of iron in the steel.

6. The reactions involved in the manufacture of nitric acid can be represented as follows:



Calculate:

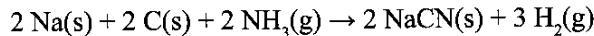
- (a) the mass of nitrogen, and
(b) the volume of nitrogen at S.T.P. required to manufacture 25.0 kg of nitric acid.

7. Methane (CH_4) burns in oxygen to form carbon dioxide and water. Write a balanced equation for the process.

If 2.50 kg of water is produced at S.T.P., calculate the volumes of:

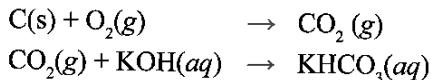
- (a) carbon dioxide produced, and
(b) methane consumed at the same temperature and pressure.

8. Sodium cyanide, which is used in the extraction of gold, can be made from sodium metal, carbon and ammonia:



Calculate the mass of sodium cyanide that would be formed by reacting 0.500 tonne of sodium metal with 762 kL of ammonia gas at S.T.P. in the presence of excess carbon.

9. It has been suggested that methane gas be used as an alternative fuel to petrol (assume C_8H_{18}) in motor vehicles, as it produces less of the greenhouse gas, $CO_2(g)$, for the same energy output.
- Write separate equations for the combustion of each fuel, methane (CH_4) and petrol (C_8H_{18}).
 - The combustion of 206 g of methane produces as much energy as 239 g of petrol. Determine the volume of $CO_2(g)$ produced at STP, for each mass of fuel.
 - Determine the percentage change in the volume of $CO_2(g)$ emission when methane is used as an alternative fuel to petrol.
 - Using your results from part (c), does methane produce less CO_2 greenhouse gas emission as claimed?
10. A 0.7941 g sample of cast iron was heated in oxygen to convert the carbon it contained to carbon dioxide. The carbon dioxide produced was absorbed in potassium hydroxide solution to form potassium hydrogencarbonate.



The potassium hydroxide solution increased in mass by 0.0732 g.
Calculate:

- The number of moles of carbon dioxide gas that dissolved in the potassium hydroxide solution.
- The volume of carbon dioxide produced at STP.
- The percentage of carbon in the cast iron.

Set 31: Solutions

Set 31: Exercises

- A student recorded the solubility of sugar at different temperatures.
 - Draw a solubility curve for sugar with temperature on the horizontal axis.
 - Describe the pattern shown by the graph.
 - Determine the solubility of sugar at:
 - 30°C
 - 70°C
 - Use the information in the student's table to describe an unsaturated, saturated and a super saturated solution of sugar at 20°C ?
 - Classify the following sugar solutions as unsaturated, saturated or super-saturated.
 - 200 g of sugar dissolved in 100 g of water at 40°C
 - 200 g of sugar dissolved in 50 g of water at 60°C
 - 50 g of sugar dissolved in 20 g of water at 100°C
- A 2.000 kg (approximately 2.0 L) sample of recycled water on a property was examined for total dissolved solids (TDS). The purpose was to determine if after treatment, its TDS concentration was at an acceptable level for use in irrigation. The sample was evaporated to dryness and the total dissolved solids remaining weighed 3.45 g.
Parts per million can be calculated as follows:

$$\text{ppm} = \frac{\text{mg of solute}}{\text{kg of solution}}$$
- Water up to 2,500 ppm can be used for irrigation. Use the information in Figure 31.1 to classify the salinity of the sample and state if it is suitable for irrigation.
- Seawater contains about 35,000 ppm of salt. How many grams of salt would be obtained if 1.000 kg (approximately 1 L) of seawater were evaporated to dryness?
- Describe the chemical tests that would allow you to distinguish between solid samples of the following barium salts: BaCO_3 , $\text{Ba}(\text{NO}_3)_2$, BaCl_2 and BaSO_4 .
- Use the solubility table (Appendix) to identify and describe any precipitate that forms when the following solutions are mixed.
 - NaCl and AgNO_3
 - $\text{Pb}(\text{NO}_3)_2$ and KI
 - K_2SO_4 and $\text{Ba}(\text{OH})_2$
 - CuSO_4 and NaOH
 - $(\text{NH}_4)_3\text{PO}_4$ and FeCl_2
- Refer to Set 27: Ionic equations. Read the explanations and examples. Write ionic equations for any precipitation reactions occurring in question 5.
- Classify the following as strong, weak or non-electrolytes: tap water, sea water, sugar solution, copper sulfate solution, hydrochloric acid solution.

Temperature $^{\circ}\text{C}$	Solubility (g/100 g water)
0	179
20	204
40	238
60	287
80	362
100	487

Fresh water	Less than 1,000 ppm
Slightly saline water	From 1,000 ppm to 3,000 ppm
Moderately saline water	From 3,000 ppm to 10,000 ppm
Highly saline water	From 10,000 ppm to 35,000 ppm

Figure 31.1

Notes

Set 32: Solution concentrations

Notes

Many reactions occur in aqueous solutions. Stoichiometry in solutions uses concentration in calculations.

Concentration is defined as the amount of solute per unit volume of solution. The amount of solute is usually expressed in moles. The relationship can be expressed as:

$$n = cV$$

where n is the number of moles of solute (mol),

c is the concentration of solute in moles per litre (mol L^{-1}), and

V is the volume of the solution in litres (L).

Note: Solution concentrations can be expressed using ' c ' or '[]' that is, $c(\text{NaCl})$ or $[\text{NaCl}]$ both refer to the concentration of NaCl in solution, expressed in mol L^{-1} .

Examples

- Find the number of moles of sodium chloride in 15.0 mL of a 2.50 mol L^{-1} NaCl solution.

$$\begin{aligned} n &= cV \\ &= (2.50)(15.0 \times 10^{-3}) \\ n &= 3.75 \times 10^{-2} \text{ mol of NaCl} \end{aligned}$$

- Calculate the concentration in mol L^{-1} of an iron(II) sulfate-7-water solution which contains 5.56 g of the salt dissolved in distilled water and made up to 750.0 mL of solution.

- Calculate the number of moles of solute:

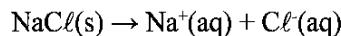
$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{5.56}{278.0} \quad M(\text{FeSO}_4 \cdot 7\text{H}_2\text{O}) = 278.0 \text{ g mol}^{-1} \\ n &= 2.00 \times 10^{-2} \text{ mol} \end{aligned}$$

- Calculate the concentration of $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ solution.

$$\begin{aligned} c &= \frac{n}{V} \\ &= \frac{2.00 \times 10^{-2}}{750 \times 10^{-3}} \\ c &= 2.67 \times 10^{-2} \text{ mol L}^{-1} \end{aligned}$$

Concentration of ions

Ionic substances and some covalent substances, such as strong acids, produce almost 100% ions when dissolved in H_2O . For example



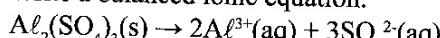
Knowing the concentration of the solute and its formula, the concentration of ions formed in solution can be calculated.

Notes

3. In 25.0 mL of a 0.200 mol L⁻¹ solution of aluminium sulfate, calculate:

(a) the $[Al^{3+}]$ (c) $n(Al^{3+})$
 (b) the $[SO_4^{2-}]$ (d) $n(SO_4^{2-})$

(i) Write a balanced ionic equation:



(ii) From the balanced ionic equation relate the unknowns to the known:

$$\begin{aligned} (a) \quad c(Al^{3+}) &= 2 \times c(Al_2(SO_4)_3) \\ &= (2)(0.200) \\ c(Al^{3+}) &= 0.400 \text{ mol L}^{-1} \end{aligned}$$

$$\begin{aligned} (b) \quad c(SO_4^{2-}) &= 3 \times c(Al_2(SO_4)_3) \\ &= (3)(0.200) \\ c(SO_4^{2-}) &= 0.600 \text{ mol L}^{-1} \end{aligned}$$

(c)(i) Calculate the number of moles of $Al_2(SO_4)_3$

$$\begin{aligned} n(Al_2(SO_4)_3) &= c(Al_2(SO_4)_3) V \\ &= (0.200)(25.0 \times 10^{-3}) \\ &= 5.00 \times 10^{-3} \text{ mol} \end{aligned}$$

(ii) Relate the unknown numbers of moles to the known number of moles of $Al_2(SO_4)_3$:

$$\begin{aligned} n(Al^{3+}) &= 2 \times n(Al_2(SO_4)_3) \\ &= (2)(5.00 \times 10^{-3}) \\ n(Al^{3+}) &= 1.00 \times 10^{-2} \text{ mol} \end{aligned}$$

$$\begin{aligned} (d) \quad n(SO_4^{2-}) &= 3 \times n(Al_2(SO_4)_3) \\ &= (3)(5.00 \times 10^{-3}) \\ n(SO_4^{2-}) &= 1.50 \times 10^{-2} \text{ mol} \end{aligned}$$

Set 32: Exercises

1. Calculate the concentrations of the following solutions:

- (a) 0.223 moles of copper(II) sulfate in 125.0 mL of solution
 (b) 1.17 moles of sodium chloride in 2.05 L of solution
 (c) 0.0335 moles of silver nitrate in 250.0 mL of solution

2. Calculate the number of moles of:

- (a) potassium nitrate in 105 mL of 2.55 mol L⁻¹ potassium nitrate solution
 (b) sodium carbonate in 2.50 L of 0.112 mol L⁻¹ of sodium carbonate solution
 (c) potassium permanganate in 660 mL of 0.230 mol L⁻¹ of potassium permanganate solution

Set 32: Solution concentrations

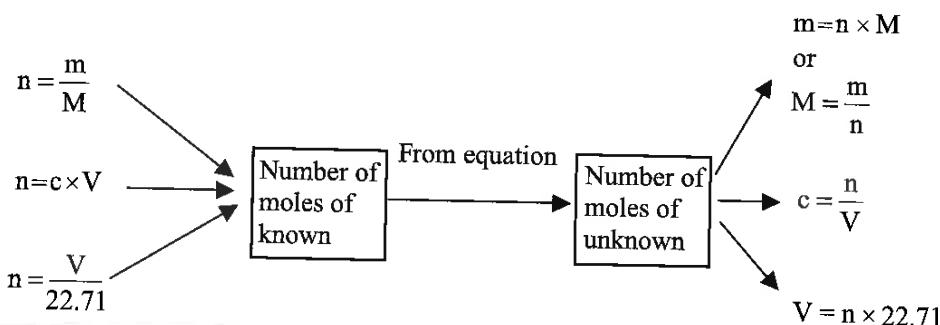
Notes

3. Calculate the mass of solute that must be used in order to prepare each of the solutions listed below:
 - (a) 630 mL of 1.26 mol L⁻¹ KCl solution from KCl
 - (b) 250 mL of 0.265 mol L⁻¹ Na₂CO₃ solution from Na₂CO₃·10H₂O
 - (c) 1.05 L of 0.420 mol L⁻¹ H₂C₂O₄ solution from H₂C₂O₄·2H₂O
4. Calculate the number of moles of:
 - (a) chloride ions in 25.0 mL of 0.200 mol L⁻¹ barium chloride solution
 - (b) sulfate ions in 550 mL of 2.56 mol L⁻¹ sodium sulfate solution
 - (c) nitrate ions in 2.20 L of 2.02×10^{-3} mol L⁻¹ lead(II) nitrate solution
5. In a solution prepared by dissolving 10.0 g of potassium carbonate in 220 mL of distilled water, determine:
 - (a) the concentration, in mol L⁻¹, of K⁺ (aq)
 - (b) the concentration, in mol L⁻¹, of CO₃²⁻ (aq)
6. Calculate the volume of 4.0 mol L⁻¹ nitric acid solution required to prepare 250 mL of 0.250 mol L⁻¹ solution.
7. Calculate the concentration in mol L⁻¹ of ammonium ions in a solution prepared by mixing 360 mL of 0.250 mol L⁻¹ ammonium sulfate solution with 675.0 mL of 1.20 mol L⁻¹ ammonium nitrate solution. Assume solution volumes are additive.
8. A laboratory assistant has a solution of 0.120 mol L⁻¹ KMnO₄. What volume of this solution must be diluted to produce 500.0 mL of a 0.025 mol L⁻¹ solution?
9. What volume of water must be added to 150.0 mL of 1.10 mol L⁻¹ sulfuric acid solution to prepare a 0.210 mol L⁻¹ solution?
10. 25.6 g of anhydrous sodium carbonate is dissolved in 200.0 mL of water.
 - (a) Calculate the concentration in mol L⁻¹ of this solution. Assume no volume change.
 - (b) A 20.0 mL sample of this solution was placed in a 150.0 mL flask and 80.0 mL of water added. Determine the concentration of sodium ions in this diluted solution.

Set 33: Reacting masses and gaseous and solution volumes

Using the stoichiometric relationship between the numbers of moles of reactants and products in a chemical reaction, a range of calculations can be carried out involving the masses, gaseous volumes and solution volumes of reactants and products. These calculations can be represented schematically as follows:

Notes



Set 33: Exercises

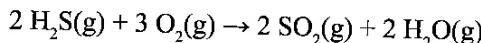
1. 25.0 mL of 0.0227 mol L⁻¹ silver nitrate solution is added to excess sodium chloride solution.
 - (a) Write an equation for the reaction.
 - (b) Calculate the mass of the precipitate formed.
2. 47.3 L of butane (C₄H₁₀) measured at S.T.P. is burned in excess oxygen.
 - (a) Write an equation for the reaction.
 - (b) What volume of carbon dioxide (measured at S.T.P.) is produced?
 - (c) What mass of water is produced?
3. 1.34 g of solid anhydrous sodium carbonate is added to excess 0.125 mol L⁻¹ hydrochloric acid solution.
 - (a) Write an equation for the reaction.
 - (b) What volume of CO₂ (measured at S.T.P.) is produced?
 - (c) What volume of hydrochloric acid would be required to react completely with the sodium carbonate?
4. Magnesium reacts with dilute hydrochloric acid according to the equation
$$\text{Mg(s)} + 2 \text{H}^+(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{H}_2(\text{g})$$
If 0.720 g of magnesium is treated with 0.950 mol L⁻¹ hydrochloric acid, calculate:
 - (a) the volume of hydrochloric acid needed to react with all the magnesium
 - (b) the volume of hydrogen gas formed at S.T.P.
 - (c) the mass of magnesium chloride which would be produced in solution
5. Hydrogen gas is produced by adding sulfuric acid to aluminium according to the equation
$$2 \text{Al(s)} + 6 \text{H}^+(\text{aq}) \rightarrow 2 \text{Al}^{3+}(\text{aq}) + 3 \text{H}_2(\text{g})$$
If 19.6 L of hydrogen is produced at S.T.P., calculate:
 - (a) the volume of 6.00 mol L⁻¹ sulfuric acid required
 - (b) the mass of aluminium consumed

Set 33: Reacting masses and gaseous and solution volumes

Notes

6. (a) What volume of 0.260 mol L^{-1} sodium iodide is needed to precipitate all the lead ions in 25.0 mL of 0.212 mol L^{-1} lead(II) nitrate?
(b) Calculate the mass of lead(II) iodide precipitated.
7. Carbon dioxide can be produced by reacting calcium carbonate with hydrochloric acid. If 10.1 g of calcium carbonate is treated with 1.07 mol L^{-1} hydrochloric acid, calculate:
(a) the volume of HCl needed to completely react with the CaCO_3 ,
(b) the volume of CO_2 produced at S.T.P.
(c) the concentration of calcium chloride in the final solution, assuming no change in the volume of solution
8. A calcite crystal (CaCO_3) with a mass of 5.60 g is placed in a beaker with 20.0 mL of dilute hydrochloric acid. The solution is warmed gently until the reaction is complete. The remaining calcite is removed and dried and found to have a mass of 5.09 g . Calculate the concentration of the hydrochloric acid.
9. An excess of hydrochloric acid was added to a sample of silver carbonate. After the reaction the silver chloride residue was washed and dried and found to have a mass of 5.74 g :
- $$\text{Ag}_2\text{CO}_3(\text{s}) + 2 \text{ HCl}(\text{aq}) \rightarrow 2 \text{ AgCl}(\text{s}) + \text{H}_2\text{O}(\ell) + \text{CO}_2(\text{g})$$
- Calculate
(a) the mass of the silver carbonate sample
(b) the number of moles of the acid consumed
(c) the volume of carbon dioxide gas, measured at S.T.P., given off during the reaction
10. When white phosphorus is heated in air it burns to form tetraphosphorus decaoxide (P_4O_{10}):
- $$\text{P}_4(\text{s}) + 5 \text{ O}_2(\text{g}) \rightarrow \text{P}_4\text{O}_{10}(\text{s})$$
- If 6.20 g of phosphorus is treated in this way:
(a) what mass of P_4O_{10} is formed?
(b) what volume of air (measured at S.T.P.) is required?
Assume that air is 20.0% oxygen by volume.
11. A 15.0 g piece of zinc is added to a beaker containing dilute hydrochloric acid and effervescence occurs:
- $$\text{Zn}(\text{s}) + 2 \text{ HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$$
- or
- $$\text{Zn}(\text{s}) + 2 \text{ H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$$
- After all reaction has ceased, the piece of zinc is washed, dried and found to have a mass of 2.00 g . Deduce from these results
(a) the number of moles of hydrochloric acid in the beaker
(b) the mass of zinc chloride which can be recovered from the solution in the beaker
(c) the volume of hydrogen gas produced by the reaction at S.T.P.

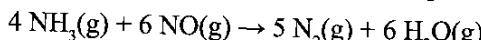
12. It has been estimated that 1.00×10^8 tonne of H_2S is released annually into the atmosphere from biogenic sources, especially from sulfate respiration by bacteria. This H_2S is oxidised to SO_2 in a day or so according to the overall process:



If 8.00 g of oxygen is consumed by a sample of hydrogen sulfide;

- (a) how many mole of hydrogen sulfide molecules would react?
- (b) what volume at S.T.P. of hydrogen sulfide would react?

13. A technique, known as the Nanaimo system for reducing the NO content of car exhausts involves the injection of a stream of NH_3 into the exhaust vapour. This converts the NO to harmless N_2 and H_2O via the reaction



Cars typically emit about 3.00 g of NO per km travelled. How many litres of NH_3 , calculated at S.T.P., would be needed to clean up the NO emitted by a car in one year, assuming that it is driven a distance of 1.60×10^4 km?

14. Liquid hydrogen is being used to power an experimental jet engine. In a particular test, in which all of the hydrogen fuel was oxidised, the fuel tank was found to have decreased in mass by 2.00 kg. What volume of oxygen, measured at S.T.P. would be needed in this test run?
15. 1.60 g of hydrogen gas is added to a container with 10.0 L of oxygen gas at S.T.P. The mixture is sparked and a reaction occurs. Find:
- (a) the mass of steam formed;
 - (b) the volume of unused oxygen at S.T.P.
 - (c) the volume of steam produced at S.T.P.
16. What volume of oxygen is required for the complete combustion of
- (a) 1.00 L of hydrogen at S.T.P.?
 - (b) 2.00×10^2 mL of methane (CH_4) at S.T.P.?
17. When 25.0 mL of a solution of dilute sulfuric acid is mixed with excess barium chloride solution, 0.483 g of precipitate is formed. What volume of $0.0134 \text{ mol L}^{-1}$ sodium hydroxide solution is required to neutralise 125 mL of the original acid?
18. A 3.00 g sample of limestone is analysed for purity by reacting it with hydrochloric acid. It is found that 20.0 mL of 2.50 mol L^{-1} HCl solution is required for complete reaction. Calculate the percentage of calcium carbonate in the limestone sample.
19. A 0.482 g sample of an alloy of silver and copper was dissolved in nitric acid. The resulting solution required 41.5 mL of $0.0993 \text{ mol L}^{-1}$ sodium chloride solution for complete precipitation of the silver ions. Calculate the percentage of silver in the alloy?

Set 34: The pH scale

The pH scale is a convenient way of expressing concentrations of hydrogen ions, $[H^+]$ in aqueous solutions. The definition of pH is expressed by the equation:

$$pH = -\log_{10} [H^+] \text{ where pH is a number without units.}$$

The pH scale has a usual range of 0 to 14 as illustrated in the diagram below:

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
$[H^+]$	$1=10^0$	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
Acidic pH < 7										Neutral pH = 7	Basic pH > 7				

Notes

Examples

- Hydrochloric acid is one of the most common industrial acids with many uses. It is produced in Western Australia in large quantities. Low concentration solutions of the acid can be standardised and used in analysis of basic solutions. Concentration of hydrogen ions is written like this $[H^+]$ using square brackets.

Calculate the $[H^+]$ and pH of a $0.0200 \text{ mol L}^{-1}$ HCl solution.

- (a) Write an ionisation equation:



- (b) Calculate the hydrogen ion concentration:

As HCl is a strong acid, it is fully ionised.

So the $[H^+] = 0.0200 \text{ mol L}^{-1}$

- (c) Calculate the pH of the solution:

$$pH = -\log_{10} [H^+] = \log_{10} (0.0200) = 1.70$$

- Orange juice has a pH of 4.50. What is the $[H^+]$ in orange juice?

Calculate the hydrogen ion concentration:

$$pH = -\log_{10} [H^+]$$

$$-4.50 = \log_{10} [H^+]$$

$$[H^+] = 10^{-4.50} = 3.16 \times 10^{-5} \text{ mol L}^{-1}$$

- A dilution Problem.

Joan, a laboratory technician, requires 250.0 mL of 0.210 mol L^{-1} hydrochloric acid solution, but finds she has only 2.00 mol L^{-1} hydrochloric acid solution. What volume of 2.00 mol L^{-1} acid should Joan use to prepare 250 mL of the new solution?

As this is a dilution problem we can use the dilution equation:
 $c_1V_1 = c_2V_2$

$$0.250 \times 0.210 = 2.00 \times V_2$$
$$\therefore V_2 = 26.3 \text{ mL}$$

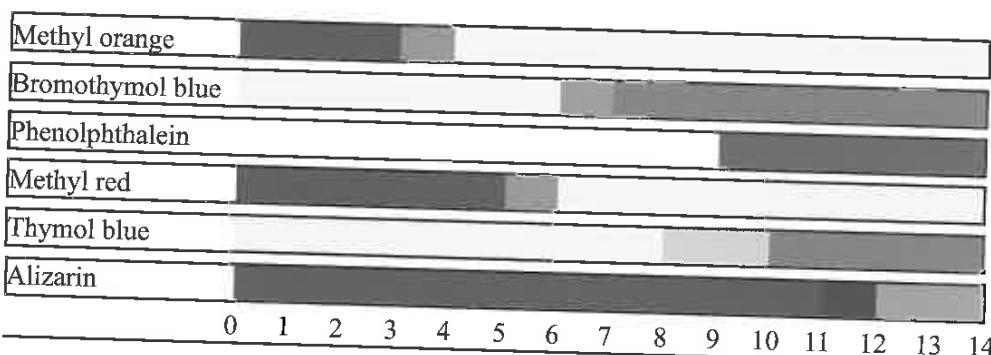
Joan should use 26.3 mL of 2.00 mol L⁻¹ acid and make it up to 250 mL with distilled water.

Notes

Set 34: Exercises

In all of the following questions the temperature is assumed to be 25 °C

Questions 1 and 2 use the following table that shows the colour ranges of some acid-base indicators.



1. A solution is yellow in methyl orange, blue in bromothymol blue and colourless in phenolphthalein. What is the pH range of the solution?
(a) 4.0 to 6.0
(b) 6.0 to 7.0
(c) 7.0 to 9.0
(d) 8.0 to 10.0
2. Mary used indicators to determine whether three colourless solutions were acidic or basic. The indicators used are shown in the table above. Samples of each solution were tested with the indicators. The colours of the resulting solutions are shown in the table.

Indicator added	Colour of Solution A	Colour of Solution B	Colour of Solution C
Methyl orange	yellow	yellow	yellow
Methyl red	yellow	yellow	yellow
Thymol blue	blue	blue	yellow
Alizarin	purple	red	red

Mary concluded that each of the three solutions tested was basic. Is she right?

3. Calculate the [H⁺] and pH of the following solutions.
(a) A 0.100 mol L⁻¹ HCl solution.
(b) A 0.00500 mol L⁻¹ HNO₃ solution.
(c) A 2.00 mol L⁻¹ HCl solution

Set 34: The pH scale

Notes

4. Calculate the $[H^+]$ of solutions with the following pH:
 - (a) Lemon juice of pH 3.00
 - (b) Dish washing solution of pH 11.0 used in a dishwasher.
 - (c) Pool acid of pH -1.00 (yes, negative one – pH exists outside 0-14)
 - (d) Orange juice of pH 4.56
 - (e) Swimming pool water of pH 7.60
5. Before disposal, a hydrochloric acid solution of pH 4.00, used to remove rust from iron parts prior to galvanising, needs to be neutralised. Base is added until the pH is 7.00. By what factor has the hydrogen ion concentration changed?
6. A brick cleaner needs to replenish the hydrochloric acid solution he is using. Using a pH meter he measures the pH of his depleted solution to be 2.00. He decides to add 3.00 L of 3.00 mol L^{-1} HCl solution to 2.00 L of his solution. Calculate the final hydrogen ion concentration of the new solution.
7. A laboratory technician wants a low concentration hydrochloric acid solution with a pH of 5.00, for calibrating a conductivity meter. What volume of distilled water must be added to a 25.0 mL sample of hydrochloric acid solution of pH 3.60 to produce the required solution?
8. A chef wishes to clean a coffee machine using citric acid ($C_6H_8O_7$). He requires 100 mL of a citric acid solution of 0.320 mol L^{-1} . Unfortunately his assistant has made up 100 mL of 0.100 mol L^{-1} solution. What mass of citric acid must he add to increase the concentration of this solution to 0.320 mol L^{-1} ? Assume the volume of the solution does not change with the addition of the solid.
9. A student prepared a 200 mL mixture made from 100 mL of HCl with pH=2 and 100 mL of HCl with pH=4. What is the new pH?
10. A water tank contains 15 000 L of bore water of pH 5.50. In an effort to increase the pH 10.0 g of sodium hydroxide was added to the water in the tank. Determine the pH of the resulting solution. Assume the volume does not change on adding the solid.
11. A swimming pool contains 2.00 ML of water at a pH of 7.80 making swimming in it unpleasant. The caretaker needs to reduce the pH to 6.80. What volume of 12.0 mol L^{-1} hydrochloric acid does he need to add? Assume the volume of acid is insignificant compared to the volume of the pool.

Set 35: Solutions of acids and bases

Moles per litre and grams per litre

Notes

Concentration is defined as the amount of solute per unit volume of solution. The units most commonly used for acid and base concentration is the mole per litre (mol L^{-1}), expressed as:

$$c = \frac{n}{V}$$

where c is the concentration of solute in moles per litre (mol L^{-1}),
 n is the number of moles of solute (mol), and
 V is the volume of the solution in litres (L)

Note: Solution concentrations can be expressed using 'c' or '[]' that is, $c(\text{NaOH})$ or $[\text{NaOH}]$ both refer to the concentration of NaOH in solution, expressed in mol L^{-1} .

Examples

- Determine the concentration in mol L^{-1} of sulfuric acid when 1.50 moles of acid is made up to 600.0 mL with distilled water.

$$c = \frac{n}{V} = \frac{1.50}{0.600} = 2.50 \text{ mol L}^{-1}$$

- Calculate the concentration in mol L^{-1} of a potassium hydroxide solution which contains 5.56 g of the solute dissolved in 750.0 mL of solution.
(a) Calculate the number of moles of solute:

$$n = \frac{m}{M} = \frac{5.56}{56.11} = 0.0991 \text{ mol} \quad M(\text{KOH}) = 56.11 \text{ g mol}^{-1}$$

- (b) Calculate the concentration of the KOH solution:

$$c = \frac{n}{V} = \frac{0.0991}{0.750} = 0.132 \text{ mol L}^{-1}$$

- Find the number of moles of sodium hydroxide in 15.0 mL of a 2.50 mol L^{-1} NaOH solution

$$n = cV = (2.50)(0.0150) = 3.75 \times 10^{-2} \text{ mol of NaOH}$$

Concentration of ions

Strong acids and strong bases produce almost 100% ions when dissolved in H_2O . Knowing the concentration of an acid or base solution and its formula, the concentration of ions present in solution can be calculated.

For example in 25.0 mL of a 0.200 mol L^{-1} solution of barium hydroxide (Ba(OH)_2) the concentrations of the individual ions are:

$$\begin{aligned}(a) c(\text{Ba}^{2+}) &= c(\text{Ba(OH)}_2) \\ &= 0.200 \text{ mol L}^{-1}\end{aligned}$$

$$\begin{aligned}(b) c(\text{OH}^-) &= 2 \times c(\text{Ba(OH)}_2) \\ &= 2 \times 0.200 \\ &= 0.400 \text{ mol L}^{-1}\end{aligned}$$

Set 35: Solutions of acids and bases

Notes

Dilution

When a solution is diluted, the volume increases and the concentration decreases. The total amount of solute, that is, the number of dissolved moles, remains the same. This can be expressed as:

$$\begin{array}{rcl} n_1 \text{ (before dilution)} & = & n_2 \text{ (after dilution)} \\ c_1 V_1 & = & c_2 V_2 \end{array}$$

Example

To prepare 250.0 mL of 0.200 mol L⁻¹ nitric acid, what volume of concentrated 15.6 mol L⁻¹ nitric acid would be required?

(a) Calculate the known number of moles:

$$\begin{array}{rcl} n(\text{dilute HNO}_3) & = & c(\text{HNO}_3)(V) \\ & = & (0.200)(250.0 \times 10^{-3}) \\ & = & 5.00 \times 10^{-2} \text{ mol} \end{array}$$

(b) Relate the unknown number of moles of concentrated HNO₃ to the known number of moles of dilute HNO₃:

$$\begin{array}{rcl} n(\text{concentrated HNO}_3) & = & n(\text{dilute HNO}_3) \\ (15.6)(V) & = & 5.00 \times 10^{-2} \\ V & = & 5.00 \times 10^{-2} \\ & & \frac{15.6}{3.21 \times 10^{-3} \text{ L}} \end{array}$$

Set 35: Exercises

1. 25.6 g of anhydrous sodium carbonate is dissolved in 200.0 mL of water.
 - (a) Calculate the concentration in grams per litre of this solution. Assume no volume change.
 - (b) Calculate the concentration in mol L⁻¹.
 - (c) A 20.0 mL sample of this solution was placed in a 150.0 mL flask and 80.0 mL of distilled water was added. Determine the concentration of sodium ions in this diluted solution.
2. Calculate the concentration of
 - (a) chloride ions in 25.0 mL of 0.200 mol L⁻¹ hydrochloric acid solution
 - (b) hydroxide ions in 2.20 L of 2.02 × 10⁻³ mol L⁻¹ barium hydroxide solution
3. 10.0 g of calcium nitrate is dissolved in 220.0 mL of distilled water. Calculate the concentration in mol L⁻¹ of:
 - (a) Ca²⁺ ions
 - (b) NO₃⁻ ions
4. Calculate the volume of 10.0 mol L⁻¹ nitric acid solution required to prepare 500.0 mL of 0.500 mol L⁻¹ solution.
5. Calculate the concentration in mol L⁻¹ of hydroxide ions in a solution prepared by mixing 360.0 mL of 0.250 mol L⁻¹ sodium hydroxide solution with 675.0 mL of 1.20 mol L⁻¹ potassium hydroxide solution. Assume solution volumes are additive.
6. What volume of water must be added to 150.0 mL of 1.10 mol L⁻¹ sulfuric acid solution to prepare a 0.210 mol L⁻¹ solution?

7. Battery acid is made by diluting (98.0% by mass) sulfuric acid with water. If 1.00 kg of the 98.0% acid was mixed with water to make up 3.00 L of battery acid, calculate the concentration of the resulting acid in g L^{-1} .
8. A 0.100 mol L^{-1} solution of hydrochloric acid was required to analyse some sandstone. Calculate the mass of hydrogen chloride gas required to make 300 mL of this solution.
9. Slaked lime (calcium hydroxide) can be used to reduce soil acidity and is sparingly soluble in water. Calculate the concentration in mol L^{-1} of:
- calcium ions;
 - hydroxide ions;
- in a solution containing 10.0 mg of calcium hydroxide in 1.00 L of solution.
10. Where vegetables are grown hydroponically, fertilisers are applied in solution. Ammonium nitrate, potassium sulfate and calcium dihydrogenphosphate are typical compounds used.
- Write equations to illustrate how each of these compounds result from acid-base reactions.
 - Calculate the mass of solute required to prepare each of the solutions listed:
 - 10.0 L of 1.50 mol L^{-1} ammonium nitrate
 - 2.50 L of 2.80 mol L^{-1} potassium sulfate
 - 500.0 mL of 0.100 mol L^{-1} calcium dihydrogenphosphate
11. Concentrated sulfuric acid is to be diluted before it is used to dissolve a sample of iron prior to analysis. Calculate the volume of 18.0 mol L^{-1} sulfuric acid required to prepare 800.0 mL of a 2.50 mol L^{-1} solution.
12. A mixture of hydrochloric acid and nitric acid is used to digest meat allowing analysis for lead, copper, cadmium and zinc. Calculate the concentration of hydrogen ions, in mol L^{-1} , in such a solution prepared by mixing 3.60×10^2 mL of a 14.3 mol L^{-1} nitric acid solution with 675 mL of 12.1 mol L^{-1} hydrochloric acid solution. Assume the solution volumes are additive.
13. Bore water with a hydrogen ion concentration of $3.60 \times 10^{-4} \text{ mol L}^{-1}$ can be relatively acidic. This can make it unsuitable for irrigation. Mixing this water with rain water reduces the concentration of hydrogen ions and hence increases the pH.
- Assuming the rain water is neutral, calculate the volume of rain water that should be added to 3.00×10^3 L of bore water to reduce the hydrogen ion concentration to $1.00 \times 10^{-6} \text{ mol L}^{-1}$.
 - Is this a practical method of reducing the hydrogen ion concentration? Explain.
 - Suggest other methods that could be used to achieve the same result. How does this compare with the addition of rain water?
14. A dilute sulfuric acid solution can be used to dissolve magnesite samples prior to analysis. Magnesite contains mostly magnesium carbonate. Calculate the volume of water that must be added to 1.50×10^2 mL of a 5.50 mol L^{-1} sulfuric acid solution to prepare a 0.500 mol L^{-1} solution for use in the analysis.



Question 8: Sandstone in hydrochloric acid



Question 10: Hydroponics (photo:
Joel Malcolm)

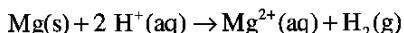
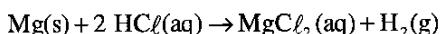
Set 36: Acid and base reaction stoichiometry

Notes

Example (mass to volume calculation)

A student carried out a laboratory experiment hoping to produce enough hydrogen gas to fill a 2.00 L balloon at S.T.P. The student reacted 3.00 g of magnesium in excess 2.0 mol L⁻¹ hydrochloric acid solution. Will enough hydrogen be produced?

- (a) Write an equation for the reaction



- (b) Calculate the number of moles of magnesium

$$n(\text{Mg}) = \frac{m}{M} = \frac{3.00}{24.3} = 0.123 \text{ mol}$$

- (c) Calculate the number of moles of hydrogen gas produced

$$n(\text{H}_2) = n(\text{Mg}) = 0.123 \text{ mol}$$

- (d) Calculate the volume of hydrogen gas produced (1 mole of a gas occupies 22.71 L at STP)

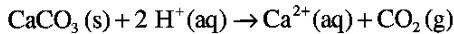
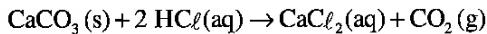
$$V(\text{H}_2) = n(\text{H}_2) \times 22.71 = (0.123) \times 22.71 = 2.79 \text{ L}$$

The reaction will produce enough gas to fill the 2.00 L balloon.

Example (volume to moles and mass calculation)

Brick cleaning acid reacts with limestone (calcium carbonate) producing carbon dioxide gas. How many moles and what mass of limestone is needed to produce 2.00 L of carbon dioxide gas at STP?

- (a) Write an equation for the reaction



- (b) Calculate the number of moles of carbon dioxide gas (1 mole of a gas occupies 22.71 L at S.T.P.)

$$n(\text{CO}_2) = \frac{V(\text{CO}_2)}{22.71} = \frac{2}{22.71} = 0.0881 \text{ mol}$$

- (c) Calculate the number of moles of calcium carbonate needed

$$n(\text{CaCO}_3) = n(\text{CO}_2) = 0.0881 \text{ mol}$$

- (d) Calculate the mass of calcium carbonate needed

$$m(\text{CaCO}_3) = n(\text{CaCO}_3) \times M(\text{CaCO}_3) = 0.0881 \times 100.09 = 8.81 \text{ g}$$

Set 36: Exercises

1. A 20.0 mL sample of stomach juices obtained from a patient suffering from hyperacidity was neutralised using 10.4 mg of solid aluminium hydroxide. Assuming that the volume of stomach juice present in the patient's stomach at the time was 250.0 mL and the only acid present was hydrochloric acid, calculate the mass of acid present in the stomach.
2. Sweat is a solution containing mainly sodium, potassium and chloride ions. During vigorous exercise, particularly in hot and humid conditions this solution often runs off the skin. In doing so it dissolves significant amounts of hydrogen ions from the skin. The concentration of hydrogen ions in sweat is typically 1.00×10^{-5} g L⁻¹. Calculate the mass of solid calcium hydroxide required to neutralise 10.0 mL of sweat.
3. A 2.89 g sample of sandstone, containing only calcium carbonate and silicon dioxide, is analysed by reacting it with hydrochloric acid. 10.7 mL of 2.50 mol L⁻¹ hydrochloric acid solution is required for complete reaction. Calculate the percentage by mass of calcium carbonate in the sandstone sample.
4. Hydrochloric acid is used to decrease the pH of swimming pool water. To determine the amount of acid to be added, the concentration of the acid needs to be known. To determine the concentration of some hydrochloric acid, 5.60 g of pure calcium carbonate in the form of calcite was placed into 20.0 mL of the acid. When the reaction was completed the remaining calcite is removed, washed in distilled water and dried. It was found to have a mass of 5.09 g. Calculate the concentration of the hydrochloric acid.
5. To determine the purity of a sample of cave limestone (mostly calcium carbonate) an analyst measured the mass of a small sample of limestone to be 2.59 g. He placed enough hydrochloric acid into a beaker to react with the limestone sample. The beaker and acid had a mass of 110.61 g. He then placed the limestone sample into the acid and allowed it to completely react. After the reaction was completed the mass of the beaker and its contents was found to be 112.22 g. Calculate the percentage purity of the limestone.
6. Sodium carbonate can be used to neutralise acid spills in the environment. It can be used as a solid or as a solution that can gradually seep into the soil.
 - (a) Calculate the concentration, in mol L⁻¹, of a sodium carbonate solution produced from 2.50×10^2 g of sodium carbonate-10-water dissolved into 2.00 L of water to produce 2.10 L of solution.
 - (b) If the acid spilt was hydrochloric acid, calculate the mass of sodium chloride produced in the soil if all of the 2.10 L of the sodium carbonate solution reacted with acid in the soil.
7. Slaked lime (Ca(OH)_2) is sometimes used to reduce soil acidity. To neutralise the top 200 mm of soil in a paddock required 1.00×10^2 g of slaked lime per square metre. Calculate the number of moles of hydrogen ions per square metre in the top 2.00×10^2 mm of soil.
8. 9.00×10^4 L of rain water with an average hydrogen ion concentration of 1.00×10^{-4} g L⁻¹ was collected and stored in a galvanised iron tank. After several months the pH of the water was found to be 7.00 ($[\text{H}^+] = 1.00 \times 10^{-7}$ mol L⁻¹). Assuming the zinc was the only material in contact with the water, calculate the
 - (a) mass of zinc that would have dissolved
 - (b) the concentration of zinc ion, in mol L⁻¹, in the rain water.



Question 4: Hydrochloric acid

Notes

Set 37: Aqueous solutions of acids and bases

Set 37: Exercises



Question 2: Marble statue



Question 5: Mortar



Question 6/7: Rust and rust remover

Double arrows \rightleftharpoons are used when writing equilibrium equations such as the formation of weak acids and bases. They indicate that both the forward and reverse reactions are occurring.

1. Write an equation to show how sulfur dioxide in the atmosphere can produce acid rain.
2. In Europe acid rain destroys the surface of historic marble structures. Marble is almost pure calcium carbonate. Write an equation to illustrate the reaction responsible for this process.
3. Fish are very sensitive to aluminium ions in their aquatic environments. Acid rain causes insoluble aluminium compounds such as aluminium oxide in the soil to dissolve, releasing aluminium ions into ground water as it moves through the soil. This water can end up in streams and lakes where fish populations may die. Write an equation to represent the reaction of aluminium oxide with acid rain.
4. Even in places where no atmospheric sulphur or nitrogen oxides are present, rainwater is still acidic. Explain why this is the case and write an equation to show how the rainwater becomes acidic.
5. Solid concrete and mortar are complex mixtures of hydroxides and carbonates that bind sand and stones together. Typical hydroxides and carbonates are those of calcium and magnesium. Concentrated hydrochloric acid is used to remove mortar that has been accidentally spilt onto brick walls by bricklayers at the time of building. Write equations to show how the acid is able to dissolve these hydroxides and carbonates.
6. Iron objects tend to rust when exposed to moist air. Rust is essentially iron(III) oxide containing some water of crystallisation. Hydrochloric acid is often used to remove the rust. Write an equation to represent this process
7. Rust Buster® is a colourless solution consisting of 30% phosphoric acid, H_3PO_4 , in water. It is used to remove surface rust from iron.
 - (a) Write three equations to show how phosphoric acid produces an acid solution.
 - (b) Comment on the extent to which each part of the process proceeds towards products.
 - (c) Write an equation to show how phosphoric acid removes rust, essentially iron(III) oxide, from the surface of iron.
8. When pure sulfuric acid is mixed with distilled water to make battery acid the process is probably best viewed as a reaction of the sulfuric acid with water to produce hydronium ions. Write equations for this process.
9. Acid soils can be a problem when used to grow certain crops. Suggest how the acidity of the soil can be reduced using naturally occurring materials. Write an equation to show how this can reduce the soil acidity.

Research**Notes**

1. Acid rain can be produced from atmospheric pollutants and atmospheric carbon dioxide and also from naturally produced organic compounds released into the air by plants. Research the process where acid rain is produced from naturally produced organic compounds. Find out what type of acids are found in this type of acid rain and discuss whether this type of acid rain is likely to be a significant problem.
2. Discuss the role of rainwater and carbon dioxide in the formation of caves in limestone rocks such as in the southwest of Western Australia and the Nullarbor Plain.
3. Research the composition of the digestive juices from your stomach and discuss the use of antacid medications.
4. Research the pH of your skin. What role does this have in maintaining a healthy skin and ultimately what is its effect on your health?
5. Extremophiles are organisms that live in conditions that would not support most living organisms. Two such groups of organisms are acidophiles and alkaliphiles. Research these two types of extremophiles.

Written answer question

Where applicable, use equations, diagrams and illustrative examples of the chemistry you are describing. Your answer should focus on the relevant chemical content specific to the situations described. Present your answer in a logical and coherent manner.

1. Hydrochloric acid versus phosphoric acid

Concentrated 12.0 mol L^{-1} hydrochloric acid has a hydrogen ion concentration of 12.0 mol L^{-1} whereas 12.0 mol L^{-1} phosphoric acid has a hydrogen ion concentration of around 0.008 mol L^{-1} . This difference in hydrogen ion concentration accounts for some of the differences in their chemical properties and hence the uses to which these acids are put.

Some uses are similar, for example both acids are used to remove rust from the surface of iron. Hydrochloric acid is used to remove rust from iron in preparation for galvanising and phosphoric acid is used to remove rust from iron prior to painting. If the iron treated with hydrochloric acid is washed with water and left in air even for a short time it rapidly turns brown. If iron treated with phosphoric acid is left in air there is no immediate appearance of any brown material on its surface. About a day after it dries, a dark coloured coating becomes apparent.

A second example of where hydrochloric acid is used extensively is to remove hardened mortar from bricks. Hardened mortar consists of sand stuck together mostly by crystals of calcium and magnesium hydroxides and some calcium and magnesium carbonates. The cleaning process involves application of concentrated hydrochloric acid, then washing it off with water.

Phosphoric acid cannot be used for removing hardened mortar from bricks.

- (a) Explain fully why the hydrogen ion concentration of 12.0 mol L^{-1} hydrochloric acid is different to the hydrogen ion concentration of 12.0 mol L^{-1} phosphoric acid.
- (b) Fully discuss the chemistry involved in the two examples of the uses of these acids described above. Your discussion should include an explanation of the similarities and differences in the chemistry of each acid when used to remove rust, and an explanation of why hydrochloric acid can be used to remove hardened mortar while phosphoric acid cannot.

Set 38: Energy changes

Set 38: Exercises

Notes

1. Identify the following reactions as endothermic or exothermic
 - (a) combustion
 - (b) respiration
 - (c) melting
 - (d) boiling
 - (e) freezing/solidification
 - (f) acid-base neutralisation
 - (g) dissociation of sodium hydroxide in water
2. Photosynthesis can be represented by the following reaction
$$6 \text{CO}_2(\text{g}) + 6 \text{H}_2\text{O}(\ell) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) + 6 \text{O}_2(\text{g}) \quad \Delta H = + 2803 \text{ kJ}$$
 - (a) How much energy is required to produce 1.00 mole of glucose?
 - (b) How much energy is required to convert 1.00 mole of carbon dioxide into glucose?
 - (c) How much energy is required to convert 45.0 g of water into glucose and oxygen?
 - (d) What mass of oxygen is produced when 1540.0 kJ of energy is consumed?
3. The pop test is used to confirm the presence of hydrogen. It is a highly exothermic reaction and can be represented by the following equation:
$$2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\ell) \quad \Delta H = -484 \text{ kJ}$$
 - (a) How much energy is produced when 1.00 mole of hydrogen is reacted with oxygen?
 - (b) How much energy is produced when 5.00 moles of water is formed?
 - (c) What mass of oxygen is required to produce 255 kJ of energy?
 - (d) Write the equation for the decomposition of water into hydrogen and oxygen and include the enthalpy change in the equation.
4. Methanol is used as a reactant in fuel cells. It can be produced using the following reaction:
$$2 \text{H}_2(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{g}) + 90 \text{ kJ}$$
 - (a) Rewrite the equation showing the heat of reaction as a ΔH .
 - (b) How much energy will be produced when 100.0 g of hydrogen is consumed?
 - (c) What mass of carbon monoxide is required to produce 25.0 kJ of energy?
 - (d) What mass of methanol will be produced when 555 kJ of energy is produced?
5. The formation of photochemical smog occurs when chemicals such as ozone and nitrogen monoxide react in the presence of sunlight.
$$\text{O}_3(\text{g}) + \text{NO}(\text{g}) \rightarrow \text{O}_2(\text{g}) + \text{NO}_2(\text{g})$$
 - (a) Ozone and oxygen are allotropes. What does this mean?
 - (b) This reaction is exothermic. Rewrite the equation to include 'heat' in the reaction.
 - (c) Given that the activation energy is 210 kJ and the heat of reaction is -200 kJ, draw a reaction profile diagram for this reaction.
6. Separately list common examples of endothermic and exothermic reactions or processes in and around the home e.g. combustion, hot packs, change of phase. Select the reaction that has the most significance to you. Describe and explain the processes involved. Include an energy profile diagram and equations.
7. Resusable heat packs usually contain sodium acetate. Describe how these heat packs work and why they are recharged by heating in boiling water. In your answer include reaction profile diagrams and equations where appropriate.

Set 39: Rates of reaction

The rates of chemical reactions can be altered in a number of ways, for example, by changing temperature, pressure (for systems involving gases), concentration, degree of sub-division or adding a catalyst.

The collision theory can be used to explain the observed impact of these changes on the rate of a reaction. The collision theory states that, in order for a successful reaction to occur, particles must collide with

- sufficient energy (the activation energy) and
- the appropriate orientation.

Figures 39.1 and 39.2 show energy profile diagrams for an exothermic and an endothermic reaction while figure 39.3 shows the distribution of molecular energies in a system.

Figure 39.1
Exothermic reaction

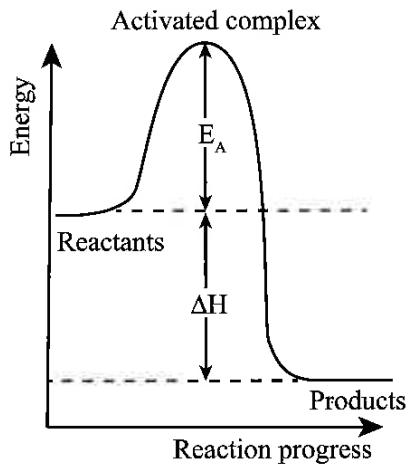


Figure 39.2
Endothermic reaction

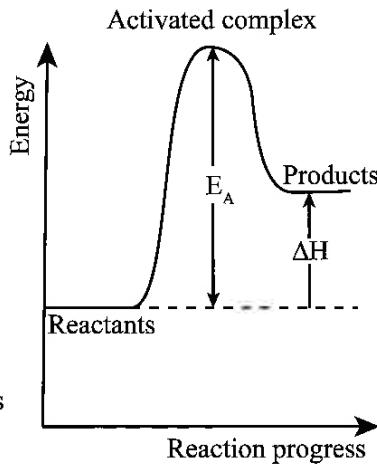
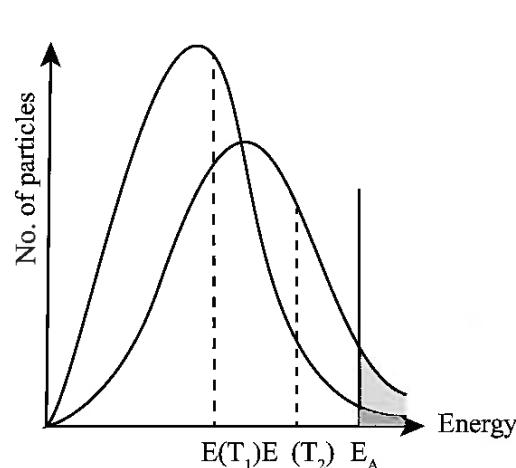


Figure 39.3
Molecular energy distribution



Exothermic reaction:

- Energy released to surroundings
- Surroundings warmer
- ΔH is negative ($-\Delta H$)

Endothermic reaction:

- Energy absorbed from surroundings
- Surroundings cooler
- ΔH is negative ($+\Delta H$)

$$E(T) = \text{average energy at temperature } T$$

$$T_1 < T_2$$

The proportion of molecules with energy equal to or greater than the activation energy, E_A (shaded areas) increases with an increase in temperature

Notes

Example

Explain why powdered calcium carbonate will react more quickly than the same mass of chips of calcium carbonate when added to dilute hydrochloric acid.

A powdered mass of calcium carbonate will have a much greater surface area than chips. A greater surface area provides more particles available for reaction. As there are more particles, there will be a greater chance of successful collision with hydrogen ions. As there are more successful collisions, there will be a greater rate of reaction.

Set 39: Exercises**Notes**

1. The following energy values were measured during a reaction.

$$E(\text{reactants}) = 110 \text{ kJ}$$

$$E(\text{products}) = 130 \text{ kJ}$$

$$E_A = 30 \text{ kJ}$$

$$\Delta H = +20 \text{ kJ}$$

(a) Draw and label an energy profile diagram for the reaction.

(b) Is the reaction described endothermic or exothermic? Explain

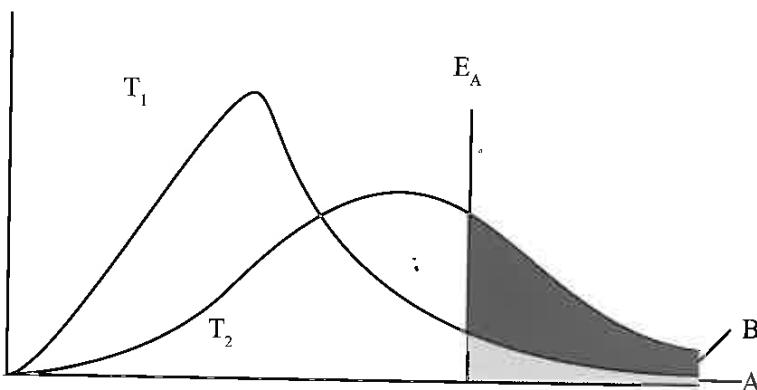
(c) This reaction is a reversible reaction. For the reverse reaction, state the

(i) ΔH

(ii) E_A

(d) On your diagram for (a), draw in a different colour and label, the reaction pathway for the reaction, if a catalyst was used.

2. The energy distribution diagram shows the distribution of kinetic energy for reacting particles at two different temperatures T_1 and T_2 . Use this diagram to answer the following questions.



(a) Label the axes.

(b) Which is the higher temperature, T_1 or T_2 ? Explain

(c) Explain what the line, E_A , represents.

(d) What does area A represent?

(e) What does area B represent?

3. Industrial processes, in order to be economic, require that products are formed as rapidly as possible at a reasonable cost. The production of ammonia through the Haber process involves an exothermic reaction between nitrogen gas and hydrogen gas to produce ammonia gas.

(a) Write a balanced equation for the reaction, including the heat of reaction.

(b) Explain how increasing the temperature will increase the rate of this reaction.

(c) Explain how increasing the pressure of the system will increase the rate of reaction.

(d) Explain how the use of a catalyst will increase the rate of reaction.

(e) Draw an energy profile diagram for the reaction.

Set 39: Rates of reaction

Notes

4. A student was experimenting with ways of increasing the reaction between marble chips (calcium carbonate) and hydrochloric acid.
 - (a) Write a balanced ionic equation for the reaction between calcium carbonate and hydrochloric acid.
 - (b) Give three ways the student could increase the rate of reaction and explain how each change causes the increase.
 - (c) What impact would using acetic acid rather than hydrochloric acid have on the rate of the reaction.
5. When carbon monoxide reacts with oxygen to produce carbon dioxide, 566 kJ of energy is produced per mole of oxygen gas consumed.
 - (a) Write a balanced equation for the reaction, including the heat of reaction.
 - (b) Draw a reaction profile diagram for this reaction given that the activation energy for the reaction is 250 kJ.
 - (c) Annotate the reaction profile diagram to show the impact of adding a catalyst to this reaction.
 - (d) List two other ways of increasing the rate of this reaction and explain how they cause the increase.
6. Flammable liquids do not burn. It is the mixture of the substance's vapour and oxygen that actually ignites. Suggest why it is the vapour that burns and not the liquid itself.
7. List five examples where rates of reaction have been altered in processes occurring in and around the home. Apply the collision theory to describe and explain each of the examples you list.
8. Enzymes are catalysts found in biological systems. Describe how enzymes act as catalysts.
9. When unleaded petrol was introduced catalytic converters were added to cars. These converters are designed to control the emissions produced in the combustion of petrol.
 - (a) Find out which chemicals are used in catalytic converters? Explain how these converters reduce the amount of noxious gases produced by cars.
Include equations in your answer.
 - (b) Leaded petrol cannot be used in cars with catalytic converters. Explain why.

Set 40: Naming and drawing hydrocarbons

Notes

The name of any simple organic compound is based on its structure. It ensures that, given the structure, a name can be constructed and if given a name, a structure can be drawn. The rules that govern this naming convention were adopted and first published by the Union of Pure and Applied Chemistry (IUPAC) in 1958. This set of rules has become known as the IUPAC Nomenclature of Organic Compounds. The rules used here include those revised in 1993.

Aliphatic hydrocarbons

For straight chain (aliphatic) hydrocarbons the rules can be summarised as follows.

1. The prefix of the name is used to identify the number of carbon atoms in a continuous chain (the chain length).

Number of carbon atoms	Prefix	Prefix for alkyl group
1	meth-	methyl
2	eth-	ethyl
3	prop-	propyl
4	but-	butyl
5	pent-	pentyl
6	hex-	hexyl
7	hept-	heptyl
8	oct-	octyl

2. The suffix of the name is used to identify the type of aliphatic hydrocarbon, that is, alkane, alkene or alkyne.

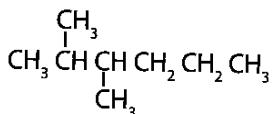
Type of hydrocarbon	Functional group	Suffix
alkane	C – C	-ane
alkene	C = C	-ene
alkyne	C ≡ C	-yne

3. Groups other than alkyl groups can result from addition and substitution reactions. The names of halogen groups are as follows.

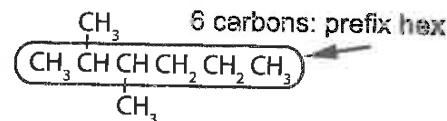
Type of substituted hydrocarbon	Functional group	Name of group
fluorocarbon	-F	fluoro-
chlorocarbon	-Cl	chloro-
bromocarbon	-Br	bromo-
iodocarbon	-I	ido-

The steps for naming hydrocarbons

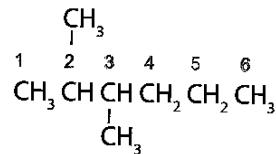
The following structure will be used to illustrate the rules for naming hydrocarbons in each step:



- Step 1: Identify the longest continuous carbon chain containing the functional group. This determines the prefix for the name.



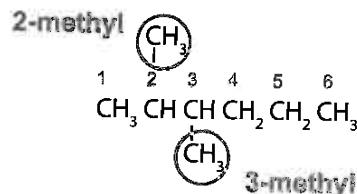
- Step 2: Assign a number to each carbon atom in the longest continuous chain starting at the end of the chain that results in the lowest possible number for the position of the functional group.



- Step 3: Identify the presence and location of any double bond. This determines the suffix of the name: -ane (only single bonds), -ene (double bond)

suffix: -ane
(only single bonds present)

- Step 4: Identify all groups attached to the longest carbon chain.

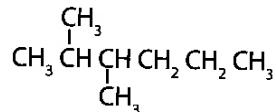


- Step 5: Add the names of the groups in front of the chain name, indicating the position of the group using the number of the carbon to which it is attached. If a group occurs more than once, use the prefix di-, tri-, tetra-, etc to indicate the number present and indicate the position of each as described above. A number for the location of groups is often used even though it may be redundant. The locating number is placed as close to the part of the name referring to the group, as in butan-1-ene.

2,3-dimethyl

- Step 6: Substituent groups are written alphabetically on the basis of the group name. Numbers are separated from each other by commas and numbers are separated from names by hyphens.

2,3-dimethylhexane



Set 40: Naming and drawing hydrocarbons

Notes

The steps for naming hydrocarbons

The following compound will be used to illustrate the rules for drawing structural formulas of hydrocarbons: 1-fluoro-2-methylbutane

Step 1: Identify and draw a carbon skeleton of the parent chain.

1-fluoro-2-methyl **butane**
C - C - C - C

Step 2: Number the carbons in the chain.

1 2 3 4
C - C - C - C

Step 3: Add the functional groups to the appropriate carbons on the chain.

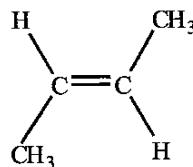
1-fluoro2-methyl butane
1 2 3 4
C - C - C - C
| |
F CH₃

Step 4: Add hydrogen atoms to give each carbon four (4) bonds.

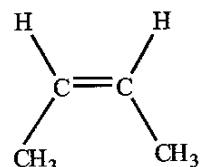
H H H H
H - C - C - C - C - H
| | |
F CH₃ H H

Cis-Trans isomers

In alkenes, rotation about the double bond is not possible. This results in the existence of cis-trans isomers in which the atoms and their bonding is the same but the arrangement of atoms in space can be different. The two different isomers are distinguished by placing the prefix *trans-* or *cis-* in front of the name, as follows



trans-but-2-ene



cis-but-2-ene

Alicyclic compounds

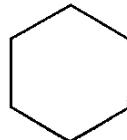
Notes

These are ring structures formed when the ends of a carbon chain join.

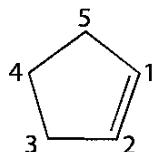
The rules for naming alicyclic (carbons forming rings) hydrocarbons are mostly the same as for those used to name aliphatic (straight chains) hydrocarbons. There are some differences though, which include the following.

1. The prefix cyclo- is placed in front of the name indicating the number of carbon atoms.
2. The carbon atoms are numbered starting at the functional group and for double bonds the numbering is such that the carbon atoms either side of the multiple bond have the numbers 1 and 2.
3. If more than one type of group is attached the numbering starts at the group that is first in alphabetical order.

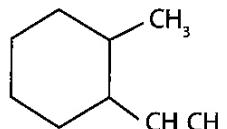
cyclohexane (C_6H_{12})



cyclopentene (C_5H_8)



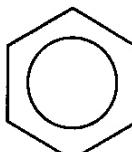
1-ethyl-2-methylcyclohexane (C_9H_{18})



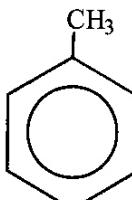
Aromatic compounds based on benzene

The naming of these compounds is very similar to the alicyclic compounds and because the compounds we will consider are based on the benzene structure, the names are based on the name benzene. This group is called 'aromatic'. For example:

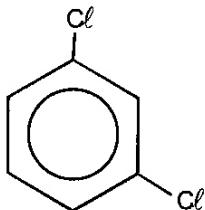
(a) benzene (C_6H_6)



(b) methylbenzene, also known as toluene (C_7H_8)



(c) 1,3-dichlorobenzene ($C_6H_4Cl_2$)



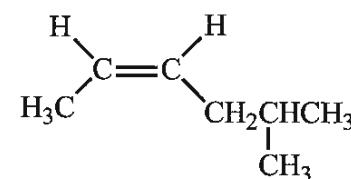
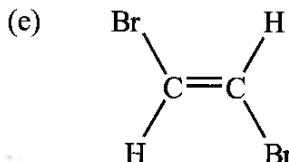
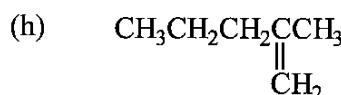
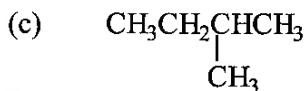
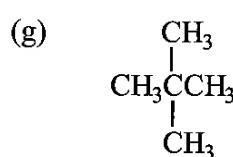
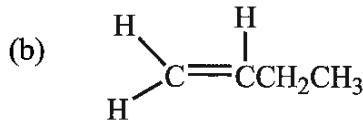
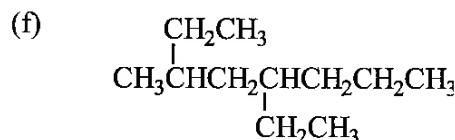
Set 40: Naming and drawing hydrocarbons

Notes

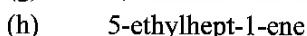
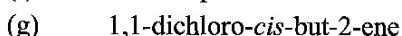
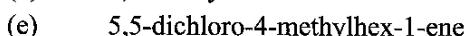
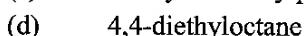
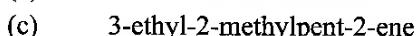
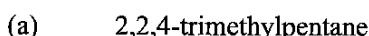
Set 40: Exercises

Petroleum is a complex mixture of hydrocarbons and can include some of the substances in **questions 1 and 2**.

1. Write systematic names for the following compounds.

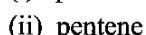


2. Draw structural formulae for the following compounds.



3. Petroleum contains many substances that have the same formula but different structures. Draw the structural isomers and write systematic names for

- (a) all the isomers of:

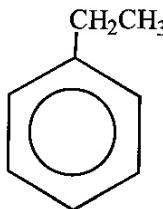


- (b) four isomers of $\text{C}_4\text{H}_9\text{Br}$

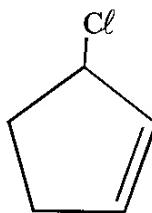
Alicyclic and aromatic compounds are often used as starting materials for important agricultural chemicals including pesticides and herbicides. Examples of substituted alicyclic and aromatic compounds are included in **questions 4 and 5**.

4. Write systematic names for the following compounds.

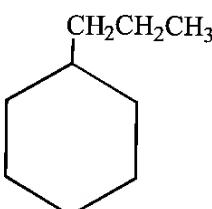
(a)



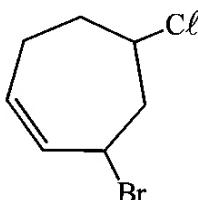
(b)



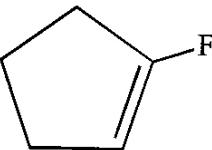
(c)



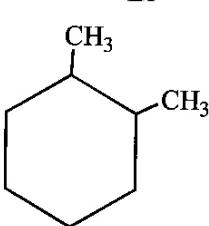
(d)



(e)



(f)



5. Draw structural formulae for the following compounds.

(a) fluorocycloheptane

(b) 4-methylcyclopentene

(c) butylbenzene

(d) 1,2-difluorobenzene

(e) 1,3-dibromobenzene

(f) 1-ethyl-4-methylbenzene

6. (a) Petroleum contains a number of different compounds with the formula C_4H_8 . Draw all the isomers of C_4H_8 .

(b) Benzene with some of its hydrogen atoms substituted with halogens can be a starting material for the synthesis of pesticides.

Draw all the isomers of dichlorobenzene.

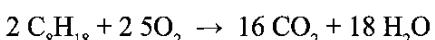
Set 41: Reactions of hydrocarbons

Notes

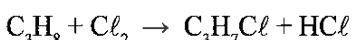
The reactions you will deal with in this section are combustion, substitution and addition reactions.

Examples

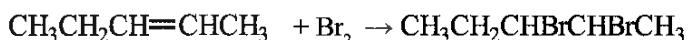
Combustion of hydrocarbons as shown in the combustion of octane



Substitution in alkanes as shown in the chlorination of propane

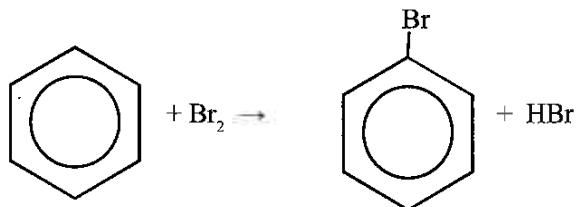


Addition in alkenes as shown in the bromination of pent-2-ene



In addition reactions involving unsymmetrical reagents such as water or hydrogen halides, Markovnikoff's rule can be applied to identify the most likely product. *Most hydrogen atoms add to the carbon atom that already has the greater number of hydrogen atoms.*

Substitution in benzene as shown in the bromination of benzene



Propane

Set 41: Exercises

1. Hydrocarbons are mostly used as fuels as they all undergo combustion to produce heat. Write equations for reactions between
 - (a) propane and oxygen when ignited.
 - (b) ethene and air when ignited.
 - (c) benzene and air when ignited
 - (d) ethylbenzene and oxygen when ignited

2. Hydrocarbons are important starting materials for the production of many useful substances. Write equations for and name any organic products formed in the reactions between
- (a) ethane and chlorine in the presence of ultraviolet radiation
 - (b) propene and bromine
 - (c) but-2-ene and hydrogen chloride
 - (d) pent-2-ene and hydrogen in the presence of a platinum catalyst
 - (e) cyclopentene and hydrogen fluoride
 - (f) benzene and bromine
 - (g) propene in excess chlorine
3. As a research chemist you are employed to investigate the conditions that give the best yield of halocarbons starting with simple hydrocarbons. You are required to make the following substances. If there is more than one type of reaction that could produce the product, use the reaction that occurs most readily.
- (i) In each case draw the structure and name the starting hydrocarbon and name any other reagents required.
 - (ii) Write equations for each step of the process.
- (a) chlorofluoromethane
 - (b) chloroethane
 - (c) 1,2-dichloroethane
 - (d) 2-chloropropane
 - (e) iodobenzene
 - (f) chlorobenzene

Set 42: Organic chemistry

Notes

Set 42: Exercises

1. Crude oil and natural gas are major sources of hydrocarbons. Research the composition of these non renewable sources of hydrocarbons, the method and extent of separation currently achieved and the uses to which these separated hydrocarbons are put.
2. One use of the hydrocarbons extracted from crude oil and natural gas is in the production of useful organic products that include solvents, refrigerants and polymers. The more useful but less common components of crude oil and natural gas are alkenes. Discuss why alkenes are more useful as starting materials than the more abundant alkanes and the role that the process of catalytic cracking has in increasing the amounts of alkenes in the products of oil refining.
3. Research the use of ethene as a raw material for the production of useful organic compounds.
4. Research the production, properties and historical uses of CFCs.



CFCs in spraycans

5. Research the impact of CFCs on atmospheric ozone. Include details of the ways in which CFCs are released into the atmosphere and where and how they interact with the ozone. You should also discuss why their interaction with ozone is a major environmental concern.
6. Research methods of removing hydrocarbons from soil contaminated with diesel or lubricating oils.
7. Currently over 90% of hydrocarbons derived from crude oil and natural gas are used as fuels, the rest are used in petrochemical industries for other uses such as solvents and raw materials for the production of other useful substances. There is a view that the use of hydrocarbons as fuels should be reduced and more hydrocarbons reserved for use as raw materials. Research arguments for and against this view.

Extended answer questions

Where applicable, use equations, diagrams and illustrative examples of the chemistry you are describing. Your answer should focus on the relevant chemical content specific to the situations described, ensure that it is presented in a logical and coherent manner.

Notes

1. Removing and preventing hydrocarbon contamination.

As an environmental scientist you are called to a truck repair workshop where local shire environmental officers have detected significant areas of the yard that have been contaminated with hydrocarbons. The main source of these hydrocarbons seems to be from the practice of washing the engines and transmissions of trucks prior to servicing and repairs, as well as washing parts prior to being repaired. The wash down area is located outside the workshop.

You are engaged to devise strategies, procedures and possibly equipment to:

- (a) remove the contaminating hydrocarbons from the soil, and
- (b) prevent further contamination while at the same time separating the hydrocarbons from the water used for washing so that it can be recycled.



2. Ground water contamination

A council gardener noticed that bore water used to reticulate gardens gradually developed a foul smell over a period of 6 months. The gardens are located approximately 100 m downhill from the local oil recycling plant that deals with used engine oil. As an environmental hydrologist you are engaged by the plant manager and the council to determine the cause of the foul smell in the water.

The manager of the recycling plant is confident that there has been no leakage from his plant and is as keen as the council to determine the source of the contamination. Your task is to determine whether or not the contamination is from the oil recycling plant.

- (a) Write possible explanations consistent with the observations.
- (b) Design an investigation or a number of investigations to test your theory or theories. Ensure that your design can be conducted safely.
- (c) What recommendations would you make to the client that could reduce or stop the problem?



Set 1: Scientific notation and unit conversions

- | | | | |
|--|---|---|--|
| 1. (a) 3.29×10^2
(e) 4×10^{-2} | (b) 1.006×10^3
(f) 7.8×10^{-8} | (c) 5.731×10^{-1} | (d) 6.724×10^{10} |
| 2. (a) 2.643×10^3 g
(e) 2.846×10^2 mg | (b) 1.74×10^{-2} g
(f) 6.72×10^2 kg | (c) 2.50×10^6 g | (d) 4.39×10^{-1} g |
| 3. (a) 1.02×10^3 cm
(e) 1.09×10^{-7} m
(i) 1.5 cm
(m) 4.9×10^{-6} m ³ | (b) 1.26×10^{-2} m
(f) 1.41×10^{-4} m ²
(j) 1 m
(n) 1.67×10^{-3} L | (c) 1.46×10^3 mm
(g) 8.3×10^{-9} m ³
(k) 1.9×10^{-4} cm | (d) 1.43267×10^2 m
(h) 1.22×10^1 m
(l) 1.5×10^3 mm ² |

Set 2: Significant figures

- | | | | |
|---|---|----------------------------|-------------------------|
| 1 (a) 3
(e) 1,2,3 or 4 | (b) 4
(f) 1 | (c) 5
(g) 3 | (d) 3
(h) 4 |
| 2 (a) 7.25 | (b) 0.0174 | (c) 6.28×10^{-3} | (d) 1.13×10^8 |
| 3 (a) 8.8 | (b) 6.60×10^2 | (c) 1.588×10^{-3} | (d) 3.34×10^4 |
| 4 (a) 6.79×10^1
(e) 2.7×10^3 | (b) 5.4×10^4
(f) 3.51×10^2 | (c) 2.83 | (d) 4.913×10^7 |
| 5 (a) 9.3 cm^3 | | | |

Set 3: Random and systematic errors

- 1 (a) 7.0 mL
(e) ± 0.25 mL
- (b) parallax error.
(c) systematic error
(d) the volume is overestimated
- 2 (a) systematic
(b) the pH readings from the meter will all be above the true value.
3. (a) parallax error if the meniscus of the liquid is always viewed either from above or below the level of the liquid.
(b) not taring (zeroing) the balance before weighing.
4. (a) differences in human perception of colour.
(b) reading an analogue scale so the uncertainty is half the smallest scale division.
(c) this relies on human perception of colour which varies with individuals.
(d) reading with a ruler so the uncertainty is half the smallest scale division.
5. Experiments that have systematic errors may have results that are not accurate. If the systematic error is consistent throughout all trials the results obtained from repeating trials can be close in value and reproducible but may not be accurate.
6. By using a digital top pan balance reading to as many decimal places as possible, e.g. using a balance that reads to 2 decimal places rather than one place will decrease the uncertainty from ± 0.1 to ± 0.01 g.
7. The reading is limited by the reaction time of the student operating the stopwatch, which is much longer than one hundredth of a second and on average is around 0.1 s.
8. (a) the results are reproducible (precise) but not accurate.
(b) the random error is relatively low as the range of values in the data set is just 0.03%. A systematic error has resulted - average of data being over 1% below the more accurate value.
(c) the accuracy of analysis is improved by altering the method of analysis not by repetition of the method.
9. (a) the mass of cracker, volume of water heated, distance burning cracker held below test tube, stirring of the test tube.
(b) energy loss to the surroundings (air, test tube) in the form of heat and light.
10. (a) heat loss to the surroundings
(b) measurement of the volumes of the solutions, the reading of the thermometer
(c) repeat the experiment several times and calculate an average.
(d) alter the method to minimise heat loss by improved insulation around cup, using a lid, and using a stirrer.
11. (a) the measurement of the masses
(b) the true value for the mass recorded could be above or below the recorded value.
(c) two measurements each with an uncertainty of ± 0.01 g are used to obtain the mass of water (13.20-12.95)
(d) the method does not heat the sample to constant mass and so not all the water has been removed.
(e) it will be an underestimate.
(f) the sample was heated to a constant mass which ensured that all of the water was removed from the hydrated sample.
(g) no, only one value is used in the calculations so the uncertainty of each measurement is irrelevant.

Answers

12. (a) $2 \text{CH}_3\text{OH} + 3 \text{O}_2 \rightarrow 2 \text{CO}_2 + 4 \text{H}_2\text{O}$ $\text{C}_2\text{H}_5\text{OH} + 3 \text{O}_2 \rightarrow 2 \text{CO}_2 + 3 \text{H}_2\text{O}$ $2 \text{C}_3\text{H}_7\text{OH} + 9 \text{O}_2 \rightarrow 6 \text{CO}_2 + 8 \text{H}_2\text{O}$
(b) ethanol 37,620 J, propan-1-ol 41,800 J
(c) ethanol 781 kJ mol⁻¹ propan-1-ol 952 kJ mol⁻¹
(d) As the molecular mass of the alcohol increases the molar heat of combustion increases.
(e) Heat loss to the surroundings including the air and the container of the water.
(f) Methanol 53 %, ethanol 57 %, propanol 47%.
(g) The sooty deposit indicates that the propan-1-ol did not combust fully into carbon dioxide and water. The soot indicates that carbon was produced.
(h) burning the fuels in oxygen enriched air to encourage complete combustion.

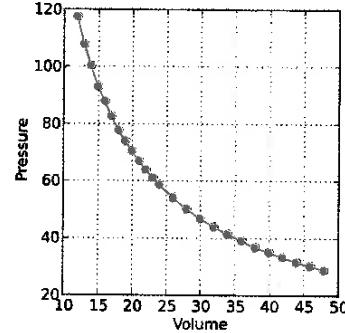
Mixtures, kinetic theory and gases

Set 4: Mixtures

1. Homogeneous mixture uniform composition and properties, heterogeneous mixture non-uniform composition and properties. Homogeneous mixtures: aqueous solution of sodium chloride, air, brass. Heterogeneous mixtures: concrete, orange juice with pulp, sandy water.
2. (a) x axis - time, y axis - temperature/°C. Freezing when the temperature is constant line is horizontal.
(b) The freezing point of pure ethanol is -114°C labelled on y axis. Pure ethanol contains only one substance - ethanol molecules.
(c) Homogeneous, water.
3. (a) The different dyes in the ink have different levels of attraction to the paper and different solubilities in water. As a result the dyes travel different distances up the paper as the solvent moves up.
(b) The ink mixture contains 3 components: a yellow, purple and blue dye.
(c) The blue dye as it has travelled the furthest along with the solvent front.
(d) An R_f value is a retardation factor or retention value. It is obtained by measuring the distance each dye has moved from where it started and dividing this by the total distance moved by the solvent.
(e) Blue dye
4. (a) Add water to dissolve sucrose, filter the insoluble sand, evaporate the water from filtrate, crystallise sucrose.
(b) Add water to dissolve sodium chloride, filter to collect the sand. Rinse with water and dry.
(c) Distillation.
(d) Paper chromatography.
(e) Fractional distillation.
5. In simple distillation boiling points of components are very different, one component with the lower boiling point is removed. Fractional distillation involves the separation of multiple components into fractions with particular boiling points.

Set 5: Kinetic theory

1. Boyle's Law: At constant temperature, the volume of a given quantity of gas is inversely proportional to its pressure: $P_1V_1 = P_2V_2$.
2. Increasing temperature increases the average kinetic energy of the particles leading to more collisions with the walls of the container resulting in a higher pressure.
3. Air particles bombard the inner wall of tyre with greater frequency and force resulting in higher pressure.
4. (a) Helium gas is less dense than air so the balloon will float/rise in air.
(b) as the balloon rises the air pressure around it decreases, the balloon expands and the pressure inside lowers
(c) as the balloon rises the external pressure continues to drop and the balloon expands so much that it bursts.
5. Boiling point occurs when the vapour pressure of the liquid being heated reaches atmospheric pressure (pressure above liquid). Atmospheric pressure decreases with altitude so the liquid can boil at a lower temperature.
6. The pressure created by vapour molecules of a substance above its liquid (or solid) form in a closed container. Solute particles interfere with the escape of solvent particles so they require more energy to escape.



Set 6: Macroscopic properties of matter

1. Supercritical carbon dioxide exists at high pressures and behaves as a gas and as a liquid. Whether a substance fits the solid/liquid/gas phase depends on the temperature and pressure it is observed at. There is also some debate about when a liquid is a liquid and when it is a solid see <http://www.newscientist.com/article/dn25441-longest-experiment-sees-pitch-drop-after-84year-wait.html#.VOv6YffnUeSo>
2. Essential to include Sodium (Na^+), Potassium (K^+), Calcium (Ca^{2+}), Magnesium (Mg^{2+}), Bicarbonate (HCO_3^-)
3. (a) Particles have more kinetic energy and collide with the walls with more energy more frequently causing more pressure rupturing the can.
(b) Air particles when cooled lose kinetic energy, will collide with less energy and less frequently resulting in less pressure. The attractive forces in the balloon overcome the air pressure and the balloon collapses. When returned to room temperature the air pressure forces the balloon to be bigger.

Answers

- 4 You might need a lot of salt! A pressure cooker is a better way to cook at higher temperature and faster. Much of pasta cooking however is water absorption, which is not so temperature dependant.

5 At higher elevations cooking times would be longer as temperature is lower

6 (a) To keep fish alive and keep an oxidising environment to prevent formation of H_2S
(b) Decreases
(c) Heating of waterways by using water to cool power stations and industrial machinery
(d) Lowers oxygen content. See part a and b

7 (a) Gas flows from high pressure to low. So it must be kept at the same pressure. Decreases volume of the gas - more portable
(b) Shallow SCUBA use only cleaned air. Deep SCUBA use helium added to their air.
(c) Nitrogen narcosis is a dulling of the senses caused by the increased solubility of nitrogen at depth (high pressure). The bends are caused when a diver surfaces quickly and the dissolved nitrogen comes out of solution and causes great pain.

8 Charles' Law describes the relationship between pressure and temperature. This can be explained using the Kinetic Theory. If the volume of a container of gas is increased at a constant temperature the particles have further to travel to strike the walls of the container. This means less frequent collisions resulting in a lower pressure.

Atomic structure and bonding

Set 7: Elements and symbols

- | | | | | |
|---|------------------|-------------------|------------------|-----------------|
| 1 | (a) fluorine F | (b) calcium Ca | (c) manganese Mn | (d) tungsten W |
| | (e) silver Ag | (f) uranium U | (g) platinum Pt | (h) iodine I |
| | (i) neon Ne | (j) ruthenium Ru | (k) thorium Th | (l) astatine At |
| | (m) germanium Ge | (n) technetium Tc | (o) barium Ba | |
| 2 | (a) Na sodium | (b) Re rhenium | (c) Cl chlorine | (d) Au gold |
| | (e) Zr zirconium | (f) Pu plutonium | (g) Ce caesium | (h) As arsenic |
| | (i) Fe iron | (j) Co cobalt | (k) P phosphorus | (I) Kr krypton |
| | (m) Zn zinc | (n) K potassium | (o) Sn tin | |

Set 8: Atoms and isotopes

1. (a) label nucleus and electron cloud (b) 3 protons, 4 neutrons and 3 electron (c) Li

2

Particle	Charge	Mass
Proton	+1	1
Neutron	0	1
Electron	-1	1/1836

3. (a) Negligible difference (b) Gives +1 charge.

4. protons, electrons, neutrons.

5. Copper: E = Cu; A = 63; Z = 29

6

Symbol	Element
^{14}C	Carbon

³⁵ Cl	Chlorine	3
------------------	----------	---

⁵⁶Fe Iron 5

Symbol	Element	A	Z	Neutrons
¹⁴ ₆ C	Carbon	14	6	8
³⁵ ₁₇ Cl	Chlorine	35	17	18
⁵⁶ ₂₆ Fe	Iron	56	26	30
³¹ ₁₅ Ga	Gallium	31	15	16
¹⁰⁸ ₄₇ Ag	Silver	108	47	61
¹² ₆ C	Carbon	12	6	6
²³ ₁₁ Na	Sodium	23	11	12
⁶⁴ ₂₉ Cu	Copper	64	29	35
⁴⁰ ₂₀ Ca	Calcium	40	20	20
¹³ ₆ C	Carbon	13	6	7

7. Carbon

8. (a) Hydrogen-1 ^1_1H (b) Hydrogen-2 ^2_1H (c) Hydrogen-3 ^3_1H

Answers

Set 9: Atomic structure and the Periodic Table

1

Z	Name	Symbol	Metal/ Non-metal	Electron configuration	Valence electron behaviour
1	Hydrogen	H	Metal	1	Loses 1, gains 1, shares 1
2	Helium	He	N/A	2	N/A
3	Lithium	Li	metal	2,1	Lose 1
4	Beryllium	Be	metal	2,2	Lose 2
5	Boron	B	Metal	2,3	Share 3
6	Carbon	C	Non-metal	2,4	Share 4
7	Nitrogen	N	Non-metal	2,5	Lose 3, share 3
8	Oxygen	O	Non-metal	2,6	Gain 2, share 2
9	Fluorine	F	Non-metal	2,7	Gain 1, share 1
10	Neon	Ne	N/A	2,8	N/A
11	Sodium	Na	metal	2,8,1	Lose 1
12	Magnesium	Mg	metal	2,8,2	Lose 2
13	Aluminium	Al	Metal	2,8,3	Lose 3
14	Silicon	Si	Non-metal	2,8,4	Share 4
15	Phosphorus	P	Non-metal	2,8,5	Lose 3, share 3
16	Sulfur	S	Non-metal	2,8,6	Gain 2, share 2
17	Chlorine	Cl	Non-metal	2,8,7	Gain 1, share 1
18	Argon	Ar	N/A	2,8,8	N/A
19	Potassium	K	metal	2,8,8,1	Lose 1
20	Calcium	Ca	metal	2,8,8,2	Lose 2

2. (a) Metals (b) Left hand side of the table

3. (a) Non-metals (b) Right hand side of the table

4. (a) Noble gases (also boron and some Group 14 elements (e.g. carbon, silicon))
(b) Far right hand side of the Table (also middle of the table)

5. The number of electrons likely to be gained or lost by an element can be predicted based on the Group the element belongs to on the Periodic Table. Elements in Groups 1 - 3 will lose electrons (the number lost will be the same as the Group number). Elements in the middle of the main Groups tend to share electrons (this only applies to the higher elements in the group) and elements in Groups 15 - 17 tend to gain electrons. Elements in Group 18 do not gain or lose electrons.

6

	Symbol	Atomic number	Number of protons	Number of electrons	Number of neutrons
Example	${}^{19}_{\text{F}}$	9	9	9	10
A	${}^{16}_{\text{O}}$	8	8	8	8
B	${}^{17}_{\text{O}}$		8	8	9
C	${}^{16}_{\text{O}}{}^{2-}$	8	8	10	8
D	${}^{17}_{\text{F}}$	9	9	9	8

7. A, B and C are isotopes and D and the example are isotopes. They have the same atomic number and different mass numbers.

8. The number of protons of an atom defines which element (identity) it is while the chemical behaviour of an atom can be explained in terms of its electron configuration (personality)

9. (a) 2,8,8 (b) 2,8 (c) 2,8,8,1 (d) 2,4

10. (a) chlorine (b) calcium (c) aluminium

11. (a) not ground state (b) not ground state (c) not ground state

Set 10: Ionisation energy

1. The energy required to remove a mol of electrons in the third shell of a mol of sodium atoms in the gaseous state

2. High energy is needed so the sample of element changes to gaseous phase before ionisation occurs.

3. On the graph the first ionisation energies of successive elements increase across a Period until Group 18 to then fall for the first element in the next Period.

4. Atoms with low first ionisation energies tend to lose electrons, while those with higher first ionisation energies tend to gain electrons.

5. The third electron is a whole shell closer to the nucleus resulting in a much higher ionisation energy than the first or second electron.

Answers

6. (a) 3 (b) A large jump in ionisation energy represents a change between electron shells.
(c) Yes, same number of valence electrons.
7. 2, a large increase in energy occurs between the second and third ionisation energies.
8. A, the large increase in ionisation energies after first ionisation energy which indicates one valence electron.
9. E, the large increase in ionisation energies occurs after the third ionisation energy.
10. Atoms with a small number of valence electrons will tend to form metallic bonds, while atoms with a larger number of valence electrons will tend to share electrons in covalent bonds.

Set 11: Periodic Trends

1. The metallic character increases.
They have the same number of electrons.
The first ionisation energy decreases.
The electronegativity decreases.
The atomic radius increases.
2. (a) Across the Period (left to right) atomic radii get smaller.
(b) As sodium loses the electron in the third energy level, so size of the ion is smaller than the atom. The chlorine gains an electron and so completes its third energy level and the size of its ion as a consequence is larger.
3. (a) The electronegativity increases as number of protons in nucleus is increasing and electrons are attracted more closely.
(b) Electronegativity increases as you move up a Group - same number of electrons but closer to nucleus.
4. It could be placed in Group 2, a metal forming strong ionic bonds with non-metals.
5. 1-3 electrons metallic; 3-4 electrons covalent network; 5-7 electrons covalent molecular.

Set 12: Properties and structures of atoms

1. (a) The different images represent the sum of knowledge to a point of time in history - they included the solid sphere, the "plum pudding" model, the Bohr model and the Quantum Mechanical model.
(b) They show increasing complexity of the structure of the atom
(c) It represents a dynamic model
(d) E,B,A,C,D, page 126
(e) Dalton, Thomson, Rutherford, Bohr, Chadwick and Planck
2. All organic materials absorb and reflect light in the near-infrared (NIR) region of the spectrum, and the particular pattern of reflectance, how much at each wavelength, is that material's spectrum. The NIR spectrum gives us a chemical 'fingerprint' of that material and this fingerprint can be used to positively identify it.
3. (a) Tc - radioactive tracer
(b) Mo - used to manufacture Tc 99
(c) Co - as a tracer and to sterilize medical equipment
4. Dalton - atomic theory Thompson - electron Rutherford - nucleus and electrons Bohr-shells Chadwick-neutron
5. To identify atoms in a sample.
6. Your diagram should include an atomiser, a radiation source, wavelength selector, detector, amplifier and signal processor

Set 13: Compounds and formulae

1. (a) carbon monoxide (b) sulfur dioxide
(c) phosphorus pentachloride (d) dinitrogen monosulfide
(e) diphosphorous tetrabromide (f) sulfur hexafluoride
2. (a) NO (b) NO₂ (c) N₂O₄
(d) SO₃ (e) H₂O (f) P₅O₁₀
(g) HCl (h) HI (i) PBr₃
(j) NH₃
3. (a) LiCl (b) AgI (c) KNO₃
(d) CsCH₃COO (e) BaBr₂ (f) CuSO₄
(g) MnO₂ (h) Ni(NO₃)₂ (i) Al₂O₃
(j) Cr₂(SO₄)₃ (k) Pb₃(PO₄)₄ (l) (NH₄)₂Cr₂O₇
4. (a) carbon monoxide (b) dinitrogen monoxide
(c) hydrogen sulfide (d) hydrogen peroxide
(e) phosphoric acid (f) calcium nitride
(g) nitric acid (hydrogen nitrate) (h) cobalt(II) hydrogenphosphate
(i) copper(I) chloride (j) iron(II) sulfate

Answers

Set 14: Bonding and properties

1. (a) covalent molecular (b) metallic
(c) ionic (d) covalent molecular
(e) ionic (f) covalent molecular
(g) metallic
2. (a) Strong electrostatic attraction between positive metal ions and delocalised electrons requires a large amount of energy to overcome the attraction and melt the solid.
(b) Delocalised electrons able to move through the lattice.
(c) Delocalised electrons are able to transfer heat energy from its source to cooler parts of the lattice.
(d) As the bonding is non-directional (delocalised electrons) layers of ions can slide over each other when a force applied.
3. Silicon dioxide is a covalent network structure. All of the bonds between the atoms are covalent bonds.
(a) Strong 3D covalent primary bonding present. Requires energy to overcome the bonding.
(b) There are no free charge carriers, so it does not conduct electricity.
(c) It is hard because it is held together by strong, 3D primary bonds.
(d) Brittle because the bonds are directional and when broken do not reform.
4. (a) Allotropes are different bonding arrangements of the same element e.g. phosphorus.
(b)
- | Allotrope | Colour | Melting point °C | Electrical conductivity | Hardness |
|------------------------|------------|------------------|-------------------------|----------|
| Diamond | Colourless | 3550 (sublimes) | Non-conductor | Hard |
| Graphite | Black | 3550 (sublimes) | Conductor | Soft |
| Fullerene (C_{60}) | Brown | 600 (sublimes) | Non-conductor | Soft |
- (c) Diamond has a 3D primary bonded structure so has a high melting point, has no free charge carriers so does not conduct, is hard due to the primary bonding between its atoms.
Graphite has a high melting point due to the extensive primary bonding between its atoms, conducts because of free electrons between the sheets and is soft because the sheets will slide over each other.
Fullerene is molecular and has weak bonds between the molecules. It is therefore soft, melts easily but does not conduct as it has no free charge carriers.
5. (a) Strong non-directional ionic bonds require high temperature to break the bonds.
(b) There are no charge carriers so it does not conduct when solid.
(c) Ions are free to move when the substance is molten and act as charge carriers.
(d) Brittle because if the ions are displaced they will be adjacent to an ion of the same charge and will repel.
(e) Sodium chloride is hard because its ions are bound together using strong primary ionic bonds.

Set 15: Uses, properties and structure

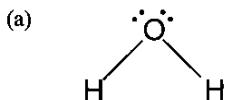
1. (a) Plastic handle (covalent molecular substance) does not conduct electricity.
(b) Plastic or wooden handle (covalent molecular substance) does not conduct heat.
(c) Conventional ovens rely on conduction. Plastic container is a poor conductor and may melt.
(d) Diamond is hard due to the strong network of covalent bonds and can be used to write on glass.
(e) Lead is cheap, low melting point for casting, dense and low reactivity.
2. Lead is very malleable, soft (and has a relatively low melting point), making it easy to mold.
3. Gold: jewellery and coinage: doesn't tarnish, very malleable and has colour. Electric circuits: excellent conductor of electricity and ductile. Shield in spacecraft: excellent reflector. Dental work: non-toxic and unreactive.
4. (a) Covalent molecular (b) Ionic
(c) Covalent molecular (d) Ionic
(e) Metallic (f) Covalent network
5. (a) Iron (b) Iron, copper(II) oxide and titanium dioxide.

Set 17: Electron dot diagrams

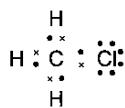
- I. (a) Na^+ (b) $\ddot{\text{Br}}^-$ (c) $\ddot{\text{P}}^-$
(d) $\left[\begin{array}{c} \text{:S:} \\ \vdots \end{array} \right]^{2-}$ (e) $\boxed{\text{H}}^+$ (f) $\left[\begin{array}{c} \text{:N:} \\ \vdots \end{array} \right]^{3-}$

Answers

2



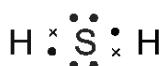
(b)



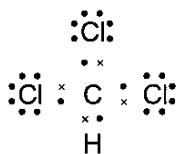
(c)



(d)



(e)

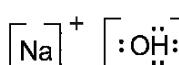


(f)

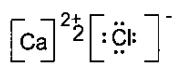


3

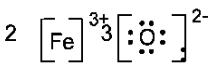
(a)



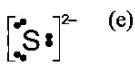
(b)



(c)



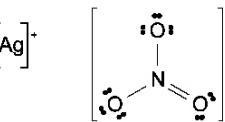
(d)



(e)



(f)



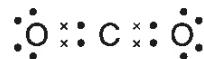
Set 18: Molecular shape

- The electronegativity of iodine is much less than for fluorine as the outer electrons of iodine are much further from the nucleus and therefore held with less attractive force compared to electrons of fluorine.
- (a) Cl e Cl Electron in the middle as each atom attracts with same force.
Cl e H Electron closer to chlorine as it has a stronger attraction.
Cl- Na⁺ Electron transferred / pulled away from sodium by chlorine.
(b) Equal sharing forms a pure covalent bond, unequal sharing forms a polar covalent bond. Transferring of an electron forms an ionic bond.
- (b), (c), (d)
-

4.

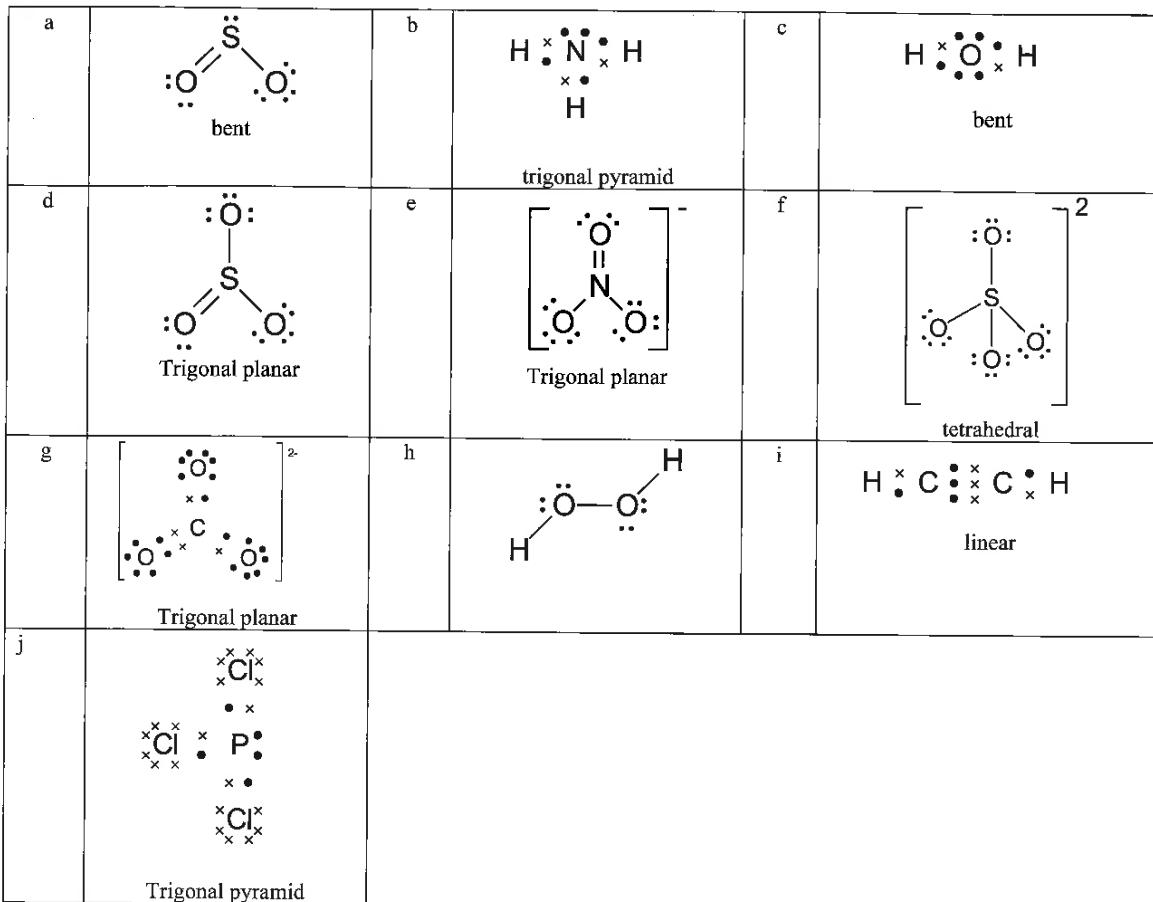
	electron dot diagram	shape	polarity
a		linear	non-polar
b		linear	non-polar
c		pyramidal	polar
d		tetrahedral	Polar
e		linear	non-polar

- The bent and pyramidal shapes will always be polar as they are asymmetrical due to the lone pair of electrons.
- The bonds within carbon dioxide for example are polar as carbon and oxygen have different electronegativities. The overall molecule is non-polar as the sum of the two dipoles cancel leaving the molecule non-polar.



Answers

7. Polarity affects the type of intermolecular force present. Non-polar molecules only have dispersion forces. Polar molecules can have dipole forces (or hydrogen bonding) as well as dispersion forces.



Set 19: Intermolecular forces I

- (a) dispersion forces
(c) hydrogen bonds and dispersion forces
(e) hydrogen bonds and dispersion forces
(g) dipole-dipole forces and dispersion forces
- (b) dispersion forces, hydrogen bonds
(d) hydrogen bonds and dispersion forces
(f) dispersion forces
(h) dispersion forces
- (a) Boiling points $\text{HCl} < \text{HBr} < \text{HI} < \text{HF}$ (b) Boiling points $\text{O}_2 < \text{CH}_3\text{Cl} < \text{NH}_3 < \text{H}_2\text{O}$
- Vapour pressure at room temperature: $\text{HNO}_3 > \text{octane} > \text{water} > \text{mercury}$
- Drawing an electron dot diagram shows if a molecule is polar and will engage in dipole-dipole bonding with water and so be soluble.
- (a) insoluble
(d) insoluble
- (b) soluble (miscible)
(e) insoluble or slightly soluble
- (c) soluble
(f) soluble
- (a) soluble
(d) soluble
- (b) insoluble
(e) soluble
- (c) insoluble
(f) insoluble

Set 20: Chromatography

- a) A liquid solvent
b) Usually a piece of glass with alumina or silica gel coating
- a) Select a beaker just taller than the TLC plate. Place a half cm of solvent in the beaker. Place a pencil line one cm from the bottom of the plate and place a drop of the sample mixture on the line. Place the end of the plate in the solvent and cover the beaker.
b) A pencil line is made of graphite, which will not contaminate the results. The dyes in ball point pen might elute.
c) Placing a lid on the beaker creates a closed system. It allows the solvent to saturate the atmosphere in the beaker and establish dynamic equilibrium. It stops the solvent evaporating.
d) Red : $2.25 \div 4.75 = 0.47$
Blue : $4.00 \div 4.75 = 0.84$
e) temperature, saturated atmosphere (do not remove the lid)
- a) Helium and dematiaceous earth with an adsorbed high boiling point liquid.
b) It might have no attraction for the stationary phase, it might have a very low boiling point.
c) The sample would be a gas and evaporate out of the liquid stationary phase quickly and then be carried on by the carrier gas.

Answers

4. a) The stationary phase can be very fine grained giving it a large surfaced area and allowing good contact between the sample and the stationary phase. This means the sample components are better separated.
b) Silica is a polar substance and the polar molecules will be attracted and have a longer retention time than less polar molecules.
5. Pressure of the solvent, temperature, surface area of the stationary phase

Chemical reactions and stoichiometry

Set 21: Relative atomic mass and mass spectroscopy

1. 35.48 2. 28.11 3. 207.24 4. Another isotope of barium with smaller atomic mass
5. 24.32 6. 72.5% of Cu-63 and 27.5% of Cu-65 7. Sample A: 60.0% of Br-79 and 40.0% of Br-81
Sample B: 5.0% of Br-79 and 95.0% of Br-81
8. (b) (i) The greater the charge the larger the deflection (ii) The larger the mass the less the deflection
(iii) I the higher charge Cu-63 is 2+ ions the greater the deflection; II the lower mass B-10 single + the greater the deflection
9. 69.60 10. 20.19 11. 39.99

Set 22: Molar mass

1. (a) 56.11 g mol⁻¹ (b) 134.5 g mol⁻¹ (c) 133.3 g mol⁻¹ (d) 74.09 g mol⁻¹
(e) 124.1 g mol⁻¹ (f) 286.1 g mol⁻¹ (g) 195.9 g mol⁻¹ (h) 16.04 g mol⁻¹
(i) 342.3 g mol⁻¹ (j) 143.3 g mol⁻¹ (k) 319.2 g mol⁻¹ (l) 142.0 g mol⁻¹
(m) 399.9 g mol⁻¹ (n) 64.06 g mol⁻¹ (o) 98.08 g mol⁻¹
2. (a) C-14 has two more neutrons in the nucleus than C-12 (b) C-13
(c) C-12 is the most abundant because the RAM of Carbon is 12.01, which is closest to the RAM of C-12 than for the other isotopes C-13 and C-14. This can be confirmed by measuring the isotopic distribution in a mass spectrometer.
3. (a) Normal hydrogen: 1 proton, 1 electron. Deuterium: 1 proton, 1 electron, 1 neutron. Tritium: 1 proton, 1 electron, 2 neutrons.
(b) Deuterium and tritium are used in nuclear research for experimental fusion reactors. Deuterated water (heavy water) is used as a moderator in some reactor designs. Tritium is used extensively in biological research as a radioactive tracer.

Set 23: Moles, particles and mass

1. (a) n = 2.96 mol (b) n = 4.00 mol (c) n = 3.99 mol
2. (a) m = 33.0 g (b) m = 1.0 × 10¹ g (c) m = 252 g
3. (a) n = 1.00 mol (b) n = 3.99 mol (c) n = 0.0500 mol
4. n = 3.50 mol 5. n = 0.116 mol
6. (a) N = 6.02 × 10²³ (b) N = 1.5 × 10²³ (c) N = 2.81 × 10²⁴ (d) N = 2.15 × 10²³
7. (a) N = 1.34 × 10²⁴ (b) N = 8.97 × 10²⁴ (c) N = 2.00 × 10¹⁹ (d) N = 2.67 × 10²³
8. (a) N = 1.48 × 10²⁴ (b) N = 1.88 × 10²³ (c) N = 3.24 × 10²¹ (d) N = 1.00 × 10²⁴

Set 24: Interpretation of formulae

1. (a) 5.00 mol (b) 0.100 mol 2. (a) 7.00 mol (b) 0.500 mol
3. (a) 160.0 g (b) 0.467 g (c) 12.0 g
4. (a) 386 g (b) 1.57 g (c) 42.8 g
5. (a) 7.50 mol (b) 2.50 mol (c) 7.50 mol (d) 30.0 mol
(e) 2.50 mol (f) 10.0 mol
6. 448 g 7. 276 g 8. 240 g 9. 600 mol nitrogen atoms 10. 1.77 g

Set 25: Percentage composition

1. (a) NaOH: Na 57.5%, O 40.0%, H 2.50%
(b) CH₃COOH: C 40.0% O 53.3% H 6.70%
(c) CuSO₄·5H₂O: Cu 25.4% S 12.9% O 57.7% H 4.00%
(d) K₃PO₄: K 55.3% P 14.6% O 30.1%
2. (a) 63.9% Cl (b) 48.0% S (c) 40.5% O (d) 3.50% N

Answers

3. (a) 62.9% (b) 41.1% (c) 14.7%
4. (a) 53.00% Bi, 32.00% Pb, 15.0% Sn (b) Bi 79.5 g, Pb 47.8 g, Sn 22.7 g
5. 7.94% Cu, 16.1% Mg, 75.9% Al 6. 80.35% Zn, 19.65% O 7. 79.9% Cu, 20.1% O
8. (a) chalcopyrite 34.6% Cu malachite 57.48% Cu (b) 2.887×10^5 g or 2.887×10^2 kg
9. Gibbsite: 34.59% Al, 34.63% H₂O Kaolinite: 20.90% Al, 13.95% H₂O
10. (a) 59.9% Ti (b) 3.12 tonnes

Set 26: Gas volumes

1. (a) 148 L (b) 19.3 L 2. (a) 0.198 mol (b) 1.10×10^{-3} mol
3. (a) 51.5 L (b) 516 L
4. (a) 1.90 L (b) 1.51 L (c) 78.1 L (d) 0.200 L
5. 94.9 g mol⁻¹ 6. 79.3 g mol⁻¹ 7. 1.41 g occupies 1L
8. (a) 8.98 g (b) 1.10×10^2 g (c) 4.45×10^{-3} g (d) 129 g

Set 27: Ionic equations

1. $\text{KC}\ell(\text{s}) \rightarrow \text{K}^+(\text{aq}) + \text{C}\ell^-(\text{aq})$ 8. $\text{Ag}^+(\text{aq}) + \text{C}\ell^-(\text{aq}) \rightarrow \text{AgC}\ell(\text{s})$
2. $\text{Ba}(\text{NO}_3)_2(\text{s}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 2 \text{NO}_3^-(\text{aq})$ 9. $\text{CO}_2(\text{g}) + \text{Ca}^{2+}(\text{aq}) + 2 \text{OH}^-(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\ell)$
3. $\text{NaOH}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$ 10. $2 \text{H}_3\text{PO}_4(\text{aq}) + 3 \text{Ag}_2\text{CO}_3(\text{s}) \rightarrow 2 \text{Ag}_3\text{PO}_4(\text{s}) + 3 \text{CO}_2(\text{g}) + 3 \text{H}_2\text{O}(\ell)$
4. $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell)$ 11. $\text{H}_2\text{S}(\text{g}) + 2 \text{Ag}^+(\text{aq}) \rightarrow \text{Ag}_2\text{S}(\text{s}) + 2 \text{H}^+(\text{aq})$
5. $\text{ZnO}(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2\text{O}(\ell)$ 12. $2 \text{H}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell) + \text{CO}_2(\text{g})$
6. $\text{CaCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$ 13. $\text{Pb}^{2+}(\text{aq}) + 2 \text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$
7. $\text{Pb}(\text{s}) + 2 \text{Ag}^+(\text{aq}) \rightarrow \text{Pb}^{2+}(\text{aq}) + 2 \text{Ag}(\text{s})$ 14. $\text{Al}(\text{OH})_3(\text{s}) + \text{OH}^-(\text{aq}) \rightarrow [\text{Al}(\text{OH})_4]^- (\text{aq})$

Set 28: Equations and observations

1. $\text{CaCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$
White solid dissolves; colourless, odourless gas evolved; colourless solution formed.
2. $\text{Mg}(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{H}_2(\text{g})$
Silver solid dissolves; colourless, odourless gas evolved; colourless solution formed.
3. $\text{NaHCO}_3(\text{s}) + 4 \text{CH}_3\text{COOH}(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$
White solid dissolves; colourless, odourless gas evolved; colourless solution formed.
4. $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell)$ No visible reaction; heat evolved.
5. $\text{H}^+(\text{aq}) + \text{NaOH}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{H}_2\text{O}(\ell)$ White solid dissolves; heat evolved.
6. $\text{CoCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{CO}_2(\text{aq}) + \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\ell)$
Pink solid dissolves; colourless, odourless gas evolved, pink solution formed.
7. $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$ White precipitate formed.
8. $\text{Pb}^{2+}(\text{aq}) + 2 \text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$ Yellow precipitate formed.
9. $2 \text{Au}^{3+}(\text{aq}) + 3 \text{Cu}(\text{s}) \rightarrow 2 \text{Au}(\text{s}) + 3 \text{Cu}^{2+}(\text{aq})$
Yellow/gold precipitate formed on surface of copper, yellow solution turns blue.
10. $2 \text{Na}(\text{s}) + 2 \text{H}_2\text{O}(\ell) \rightarrow 2 \text{Na}^+(\text{aq}) + 2 \text{OH}^-(\text{aq}) + \text{H}_2(\text{g})$
Solid reacts vigorously, may ignite; colourless, odourless gas evolved, colourless solution formed.
11. $\text{K}_2\text{CO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow 2 \text{K}^+(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$
White solid dissolves; colourless, odourless gas evolved; colourless solution formed.
12. $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell)$ No visible reaction; heat evolved.

Answers

13. $\text{Ag}^+(\text{aq}) + \text{Br}^-(\text{aq}) \rightarrow \text{AgBr}(\text{s})$ two clear and colourless solutions mixed and a cream ppt forms.
14. $\text{Ni}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Ni}^{2+}(\text{aq}) + \text{Cu}(\text{s})$ Silver metal added to blue solution. Brown/black precipitate forms solution turns green.
15. $\text{Fe}^{2+}(\text{aq}) + 2 \text{OH}^-(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s})$ a green solution and a colourless solution are mixed and a green ppt forms.

Set 29: Stoichiometry

1. (a) 4.00 mol (b) 15.0 mol (c) 3.90 mol
2. (a) $n(\text{HNO}_3) = 0.0600 \text{ mol}$ $m(\text{HNO}_3) = 3.78 \text{ g}$ (b) $n(\text{Mg}(\text{NO}_3)_2) = 0.0300 \text{ mol}$ $m(\text{Mg}(\text{NO}_3)_2) = 4.45 \text{ g}$
3. (a) 0.200 mol $m = 33.9 \text{ g}$ (b) 0.100 mol $m = 11.1 \text{ g}$ (c) 0.100 mol $m = 16.4 \text{ g}$
4. (a) m H_2SO_4 required = 1.23 g (b) Mol CuO = 0.0126 mol
5. $2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2$ 1.20 mol 6. $2 \text{HgO} \rightarrow 2 \text{Hg} + \text{O}_2$ 3.84 g
7. (a) m $\text{HNO}_3 = 3.78 \text{ g}$ (b) m $\text{CO}_2 = 1.32 \text{ g}$ (c) m $\text{Ca}(\text{NO}_3)_2 = 4.92 \text{ g}$
8. (a) m HF = 0.2222 g (b) m $\text{F}_2 = 0.1055 \text{ g}$ (c) m $\text{UF}_6 = 0.9776 \text{ g}$
9. (a) m = CaO = 258 kg (b) m $\text{CO}_2 = 202 \text{ kg}$

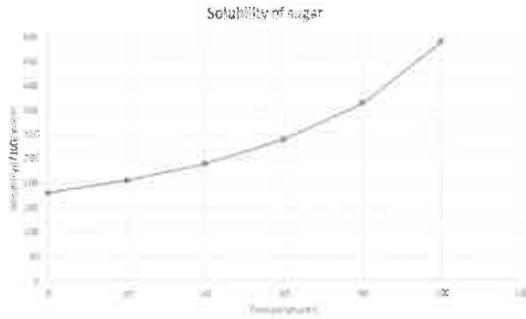
Set 30: Stoichiometry and gas volumes

1. (a) 5.68 L (b) 8.52 L (c) 0.341 L
2. 2.27 L 3. 7.81 L 4. 95.0 % pure 5. 97.0 % pure
6. (a) 5557 g (b) $4.50 \times 10^3 \text{ L}$
7. $\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$ (a) V $\text{CO}_2 = 1576 \text{ L}$ (b) V $\text{CH}_4 = 1576 \text{ L}$ 8. m NaCN = 1.06 tonnes
9. (a) $\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$ $2 \text{C}_8\text{H}_{18} + 25 \text{O}_2 \rightarrow 16 \text{CO}_2 + 18 \text{H}_2\text{O}$
 (b) $\text{CH}_4:$ V $\text{CO}_2 = 291.7 \text{ L}$ $\text{C}_8\text{H}_{18}:$ V $\text{CO}_2 = 380.3 \text{ L}$
 (c) % change: $(380.3 - 291.7) \div 380.3 = 23.3 \%$ drop in CO_2 output (d) yes
10. (a) n $\text{CO}_2 = 0.001664 \text{ mol}$ (b) V = 0.03778 L (c) 2.52 %

Solutions and acidity

Set 31: Solutions

1. (a) See graph
 (b) Sugar solubility increases rapidly with temperature
 (c) $30^\circ\text{C} = 220 \text{ g}/100\text{g water}$, $70^\circ = 320 \text{ g}/100\text{g water}$
 (d) At 20°C unsaturated < 204g/100g water
 Saturated solution = 204g/100g water,
 Supersaturated > 204g/100g water
 (e) unsaturated, supersaturated, unsaturated
2. 1725 ppm. Suitable for use
3. 35g
4. Step 1: Add a little distilled water to each – the insoluble one is barium sulphate. Step 2: Add acid to each – the carbonate will fizz.
 Step 3: Add silver nitrate solution to the remaining 2. The chloride will produce a white precipitate. Remaining sample is barium nitrate.
5. (a) silver chloride – white precipitate
 (b) lead (II) iodide – yellow precipitate
 (c) barium sulfate – white precipitate
 (d) copper hydroxide - pale blue solid
 (e) iron (III) phosphate-green precipitate forms
6. (a) $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$
 (b) $\text{Pb}^{2+}(\text{aq}) + 2 \text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$
 (c) $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
 (d) $\text{Cu}^{2+}(\text{aq}) + 2 \text{OH}^-(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$
 (e) $\text{Fe}^{2+}(\text{aq}) + \text{PO}_4^{3-}(\text{aq}) \rightarrow \text{Fe}_3(\text{PO}_4)_2(\text{s})$
7. Strong: tap water, sea water, copper sulfate, hydrochloric acid Non-electrolyte: sugar



Answers

Set 32: Solution concentrations

1. (a) 1.78 mol L^{-1} (b) 0.571 mol L^{-1} (c) 1.78 mol L^{-1}
 2. (a) 0.268 mol (b) 0.280 mol (c) 0.152 mol
 3. (a) 59.2 g (b) 19.0 g (c) 55.6 g
 4. (a) 0.0100 mol (b) 1.41 mol (c) 0.00889 mol
 5. (a) 0.658 mol L^{-1} of K^+ ion (b) 0.329 mol L^{-1} of K^+ ion
 6. 16 mL
 7. 0.956 mol L^{-1} 8. 104 mL 9. 786 mL total final volume. So we need to add $786 - 150 = 636 \text{ mL}$
 10. (a) 1.21 mol L^{-1} (b) 0.484 mol L^{-1}

Set 33: Reacting masses and gaseous and solution volumes

1. (a) $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$ (b) m AgCl formed = 0.0814 g
 2. (a) $2 \text{C}_4\text{H}_{10}(\text{g}) + 13 \text{O}_2(\text{g}) \rightarrow 8 \text{CO}_2(\text{g}) + 10 \text{H}_2\text{O}(\text{g})$ (b) 189 L (c) 188 g
 3. (a) $\text{Na}_2\text{CO}_3(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow 2 \text{NaCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$ (b) 0.287 L (c) 0.202 L
 4. (a) 0.0623 L (b) 0.672 L (c) 2.82 g
 5. (a) 0.144 L (b) 15.5 g 6. (a) 0.0408 L (b) 2.44 g
 7. (a) 0.189 L (b) 2.29 L (c) 0.534 mol L^{-1}
 8. $\text{CaCO}_3(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\ell)$ 0.509 mol L^{-1}
 9. (a) 5.52 g (b) 0.0400 mol (c) 0.454 L 10. (a) 14.2 g (b) 28.4 L
 11. (a) 0.398 mol (b) 27.1 g (c) 4.51 L 12. (a) 0.167 mol (b) 3.79 L
 13. $2.42 \times 10^4 \text{ L}$ 14. $1.12 \times 10^4 \text{ L}$ 15. (a) 14.4 g (b) 0.988 L (c) Nil
 16. (a) 0.500 L (b) 0.400 L 17. 1.54 L 18. $83.3\% \text{ CaCO}_3$ 19. $92.2\% \text{ silver}$

Set 34: The pH scale

1. (c)
 2.

Indicator added	Colour of solution A	Colour of solution B	Colour of solution C
Conclusion	$\therefore >12$	$\therefore \text{between } 10\text{-}11$	$\therefore \text{between } 6\text{-}8$
	Basic	Basic	Basic, neutral or acidic

3. (a) $[\text{H}^+] = 0.100 \text{ mol L}^{-1}$ pH = 1.00 (b) $[\text{H}^+] = 0.00500 \text{ mol L}^{-1}$ pH = 2.30
 (c) $[\text{H}^+] = 2.00 \text{ mol L}^{-1}$ pH = 0.301
 4. (a) Lemon juice $[\text{H}^+] = 1.00 \times 10^{-3} \text{ mol L}^{-1}$ (b) Dish washing solution $[\text{H}^+] = 1.00 \times 10^{-11} \text{ mol L}^{-1}$
 (c) Pool acid $[\text{H}^+] = 10.0 \text{ mol L}^{-1}$ (d) Orange juice $[\text{H}^+] = 2.75 \times 10^{-5} \text{ mol L}^{-1}$
 (e) Swimming pool water $[\text{H}^+] = 2.51 \times 10^{-8} \text{ mol L}^{-1}$
 5. Concentration changed by a factor of 1000 6. $[\text{H}^+] = 1.80 \text{ mol L}^{-1}$ 7. 603 mL 8. $4.23 \text{ g citric acid}$
 9. pH = 2.30 10. pH = 5.83 11. 5.31 mL of the hydrochloric acid solution

Set 35: Solutions of acids and bases

1. (a) 128 g L^{-1} (b) 1.21 mol L^{-1} (c) 0.484 mol L^{-1} 2. (a) 0.200 mol L^{-1} (b) $4.04 \times 10^{-3} \text{ mol L}^{-1}$
 3. (a) 0.277 mol L^{-1} (b) 0.554 mol L^{-1} 4. 0.025 L 5. 0.870 mol L^{-1} 6. 0.636 L to be added
 7. 327 g L^{-1} 8. 1.09 g 9. (a) $0.000135 \text{ mol L}^{-1}$ (b) $0.000270 \text{ mol L}^{-1}$
 10. (a) $\text{HNO}_3(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4\text{NO}_3(\text{aq}) \rightarrow \text{NH}_4\text{NO}_3(\text{s})$ (b) m(NH_4NO_3) = 1200 g
 $\text{H}_2\text{SO}_4(\text{aq}) + 2 \text{KOH}(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}(\ell)$ m(K_2SO_4) = 1220 g
 $\text{Ca}(\text{OH})_2(\text{aq}) + 2 \text{H}_3\text{PO}_4(\text{aq}) \rightarrow \text{Ca}(\text{H}_2\text{PO}_4)_2(\text{aq}) + \text{H}_2\text{O}(\ell)$ m($\text{Ca}(\text{H}_2\text{PO}_4)_2$) = 11.8 g
 11. 0.111 L 12. 12.9 mol L^{-1} 13. (a) add $1.077 \times 10^6 \text{ L}$ (b) too much rainwater (c) Add a low cost base
 14. add 1.50 L of distilled water

Answers

Set 36: Acid and base reaction stoichiometry

1. 0.182 g 2. 3.70×10^{-6} g 3. 46.3 % CaCO_3 4. 0.510 mol L⁻¹ 5. 86.1 %
6. (a) 0.4160 mol L⁻¹ (b) 102 g 7. 2.70 mol 8. (a) 291 g (b) 4.96×10^{-5} mol L⁻¹

Set 37: Acids and bases

1. $\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{SO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{HSO}_3^-(\text{aq})$
2. $\text{CaCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\ell)$
3. $\text{Al}_2\text{O}_3(\text{s}) + 6 \text{H}^+(\text{aq}) \rightarrow 2 \text{Al}^{3+}(\text{aq}) + 3 \text{H}_2\text{O}(\ell)$
4. $\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$
5. $\text{Ca(OH)}_2(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2 \text{H}_2\text{O}(\ell)$
 $\text{CaCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\ell)$
6. $\text{Fe}_2\text{O}_3(\text{s}) + 6 \text{H}^+(\text{aq}) \rightarrow 2 \text{Fe}^{3+}(\text{aq}) + 3 \text{H}_2\text{O}(\ell)$
7. $\text{H}_3\text{PO}_4(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{H}_2\text{PO}_4^-(\text{aq})$
 $\text{H}_2\text{PO}_4^-(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{HPO}_4^{2-}(\text{aq})$
 $\text{HPO}_4^{2-}(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$
8. $\text{H}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{HSO}_4^-(\text{aq})$
 $\text{HSO}_4^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$
9. Powdered CaCO_3
 $\text{CaCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\ell)$
10. (a) $\text{NaOH}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$
(b) $\text{Al(OH)}_3(\text{s}) + \text{OH}^-(\text{aq}) \rightarrow [\text{Al(OH)}_4]^- (\text{aq})$

Energy changes and rates of reaction

Set 38 Energy changes

1. Exothermic – a, b, e, f, g Endothermic – c, d
2. (a) 2.80×10^3 kJ (b) 467 kJ (c) 1.70×10^3 kJ (d) 105 g
3. (a) 242 kJ (b) 1210 kJ (c) 16.9 g (d) $2 \text{H}_2\text{O}(\ell) + 484 \text{ kJ} \rightarrow 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g})$
4. (a) $2 \text{H}_2(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{g}) \Delta H = -90 \text{ kJ}$ (b) 2232 kJ (c) 7.78 g (d) 198 g
5. (a) They are molecules representing different structural combinations of an element (b) $\text{O}_3(\text{g}) + \text{NO}(\text{g}) \rightarrow \text{O}_2(\text{g}) + \text{NO}_2(\text{g}) + \text{heat}$
6. Endothermic: boiling water, melting ice, cooking an egg and baking bread. Exothermic: rusting of iron, souring of milk, making ice blocks and burning a candle. Boiling water: When water boils energy is absorbed by the liquid, water molecules so they have enough energy to separate from the attraction for one another in the liquid to be released as gas particles.
7. A reusable sodium acetate heat pack contains a supersaturated solution of sodium acetate. The supersaturated solution is unstable, crystallizes easily and is exothermic, so the pack gets hot. To restore the pack ready to be used again, simply redissolve the crystals into the solution, by heating the pack in hot water. $\text{CH}_3\text{COONa}(\text{aq}) \rightarrow \text{CH}_3\text{COONa}(\text{s}) + \text{Heat}$

Set 39: Rates of reactions

1. (b) endothermic as the ΔH is positive (c) (i) $\Delta H = -20 \text{ kJ}$ (ii) $E_a = 10 \text{ kJ}$ (d) The ΔH is the same but the E_a is lower
2. (a) Vertical axis: Number of particles, Horizontal axis: Energy (b) $T_2 > T_1$ as it has a higher average kinetic energy
(c) E_a – the activation energy for the reaction.
(d) Area A proportion of particles with energy equal to or greater than activation energy at temperature T_1 .
(e) Area B increase in proportion of particles with energy equal to or greater than activation energy when T_1 is increased to T_2 .
3. (a) $\text{Heat} + \text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
(b) Increasing temperature increases the number of collisions and particles that have enough energy to react.
(c) Increasing the pressure increases the concentration so collisions can occur more frequently.
(d) A catalyst provides an alternative reaction pathway with a lower energy of activation.
4. (a) $\text{CaCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{H}_2\text{O}(\ell) + \text{CO}_2(\text{g})$
(b) Increase the surface area of the reaction by crushing the marble chips into smaller particles. Increase the concentration of the acid for more collisions. Heat the mixture to provide more particles with energy equal to or greater than the E_a .
(c) Decrease the rate of reaction. Acetic acid is a weaker acid so $[\text{H}^+]$ available to collide and react is less.
5. (a) $2 \text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{CO}_2(\text{g}) + \text{Heat}$
(d) Increase the pressure, which increases the concentration causing more collision and a faster reaction.
Increase the temperature to increase the number of particles with energy equal to or greater than the E_a .
6. Vapour particles collide with oxygen to react.
7. Souring of milk left out of the fridge - the higher temperature means that more milk particles have enough energy for the souring reaction to take place. Dissolving sugar or coffee in hot water - the higher temperature means that more particles have enough energy for the dissolving process to take place at a faster rate. Stirring when cooking - increases the surface area causing more collisions.
Cutting potatoes for baking or boiling - increases the surface area and enables more collisions.

Answers

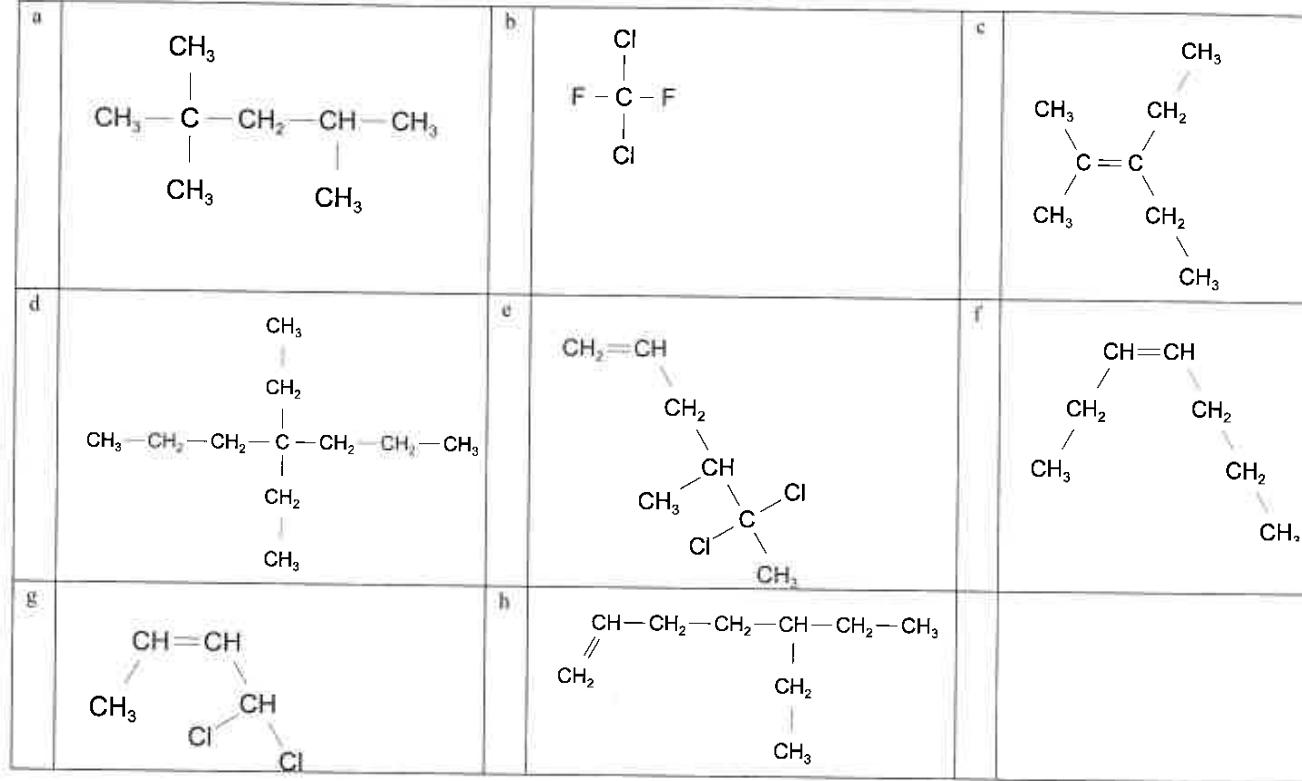
8. Enzymes work by helping with the orientation of complex particles creating more reaction collisions with suitable orientation for reactions to be successful.
9. (a) Platinum and rhodium and or palladium. $2 \text{NO} \rightarrow \text{N}_2 + \text{O}_2$ $2 \text{NO}_2 \rightarrow \text{N}_2 + 2 \text{O}_2$ $2 \text{CO} + \text{O}_2 \rightarrow 2 \text{CO}_2$
 (b) Lead in leaded fuel reacts with the metal catalyst in the converter destroying its capacity to function.

Organic chemistry

Set 40: Naming and drawing hydrocarbons

1. (a) pentane (b) but-1-ene (c) 2-methylbutane
 (d) 1,1,2-trichloroethane (e) *trans*-1,2-dibromoethene (f) 5-ethyl-3-methyloctane
 (g) 2,2-dimethylpropane (h) 2-methylpent-1-ene (i) 1,1,1-trifluoropentane
 (j) 5-methyl-*cis*-hex-2-ene

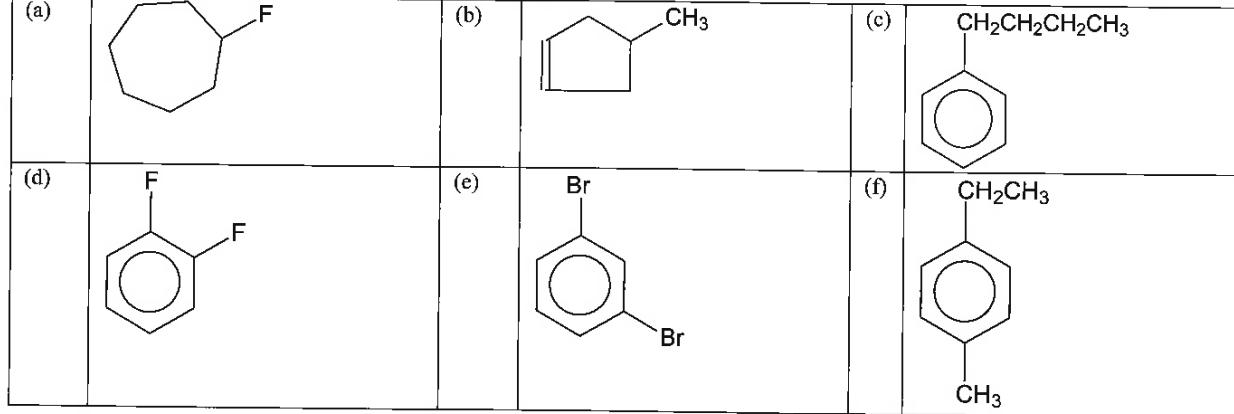
2.



3. (a) (i) Pentane, 2-methylbutane, 2,2-dimethylpropane (ii) *cis*-2-pentene and *trans*-2-pentene
 (b) 1-bromobutane 2-bromobutane 1-bromo-2-methylpropane 2-bromo-2-methylpropane

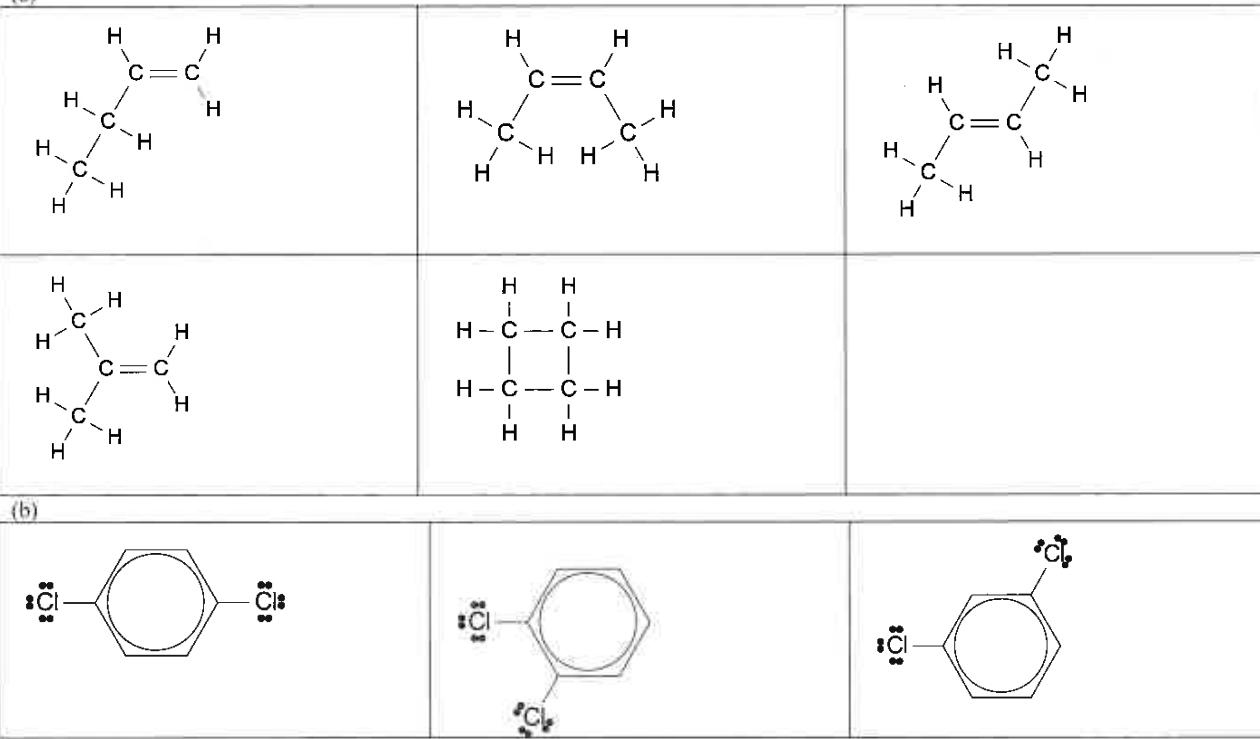
4. (a) ethylbenzene (b) 3-chlorocyclopentene
 (c) propylcyclohexane (d) 3-bromo-5-chlorocycloheptene
 (e) 1-fluorocyclopentene (f) 1,2-dimethylhexane

5.



Answers

6.



Set 41: Reactions of hydrocarbons

